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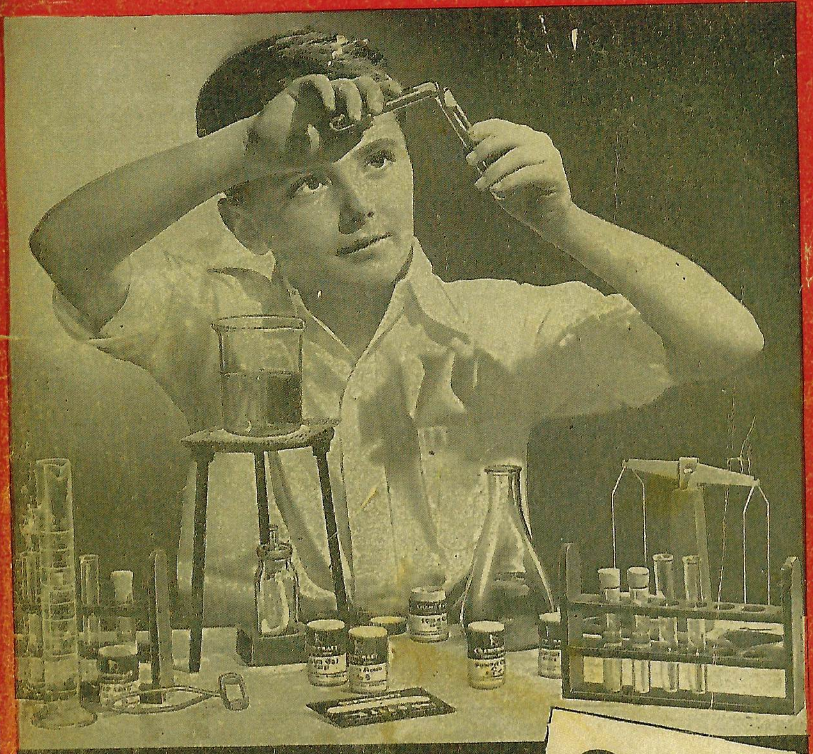
CHEMCRAFT

EXPERIMENT BOOK

INSTRUCTIVE AND FASCINATING CHEMICAL EXPERIMENTS

o Dangerous Poisons

or Explosive Chemicals



The PORTER CHEMICAL COMPANY
HAGERSTOWN • MARYLAND

Senior Manual

THE SENIOR CHEMCRAFT MANUAL

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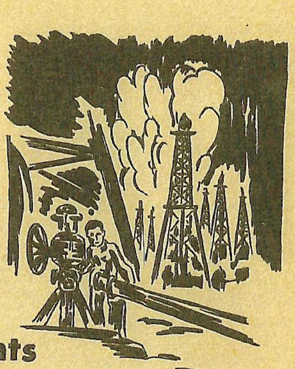

Directions for Experiments

Explanation of Chemical Science and Industries

by

Harold M. Porter

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THE PORTER CHEMICAL COMPANY
Hagerstown, Maryland

HOW TO USE YOUR CHEMCRAFT OUTFIT

CHEMCRAFT Chemistry Outfits contain no dangerous substances, but like many commonplace things they should be used in the proper way and with common sense.

By careful study and research we have prepared a great number of interesting experiments which can be performed with absolute safety if the directions in this manual are followed.

If you should eat a quantity of shaving cream it would probably make you very sick. The same is true of some of the chemicals in this outfit, although none of them are dangerous to life in the quantities supplied. If you were to strike a match while holding your finger next to the head, or ignite a whole box of matches at one time, you might be severely burned. Common sense warns you against such errors and you should exercise the same common sense when using your CHEMCRAFT. Do not combine any other chemicals, substances or materials with the chemical reagents contained in this outfit. Do not find fault with CHEMCRAFT because of incorrect results which may follow combinations of chemical compounds other than those suggested in this book. Do not expect to spill chemicals on your clothes, the furniture or the carpet without causing damage, just as happens when you spill paint, molten paraffin, milk or molasses. Always wash your hands when through experimenting, particularly after you have handled chemicals.

CHEMCRAFT should not be put into the hands of persons who are unable to read and understand the experiments in this Manual.

THE PORTER CHEMICAL COMPANY
Hagerstown, Maryland

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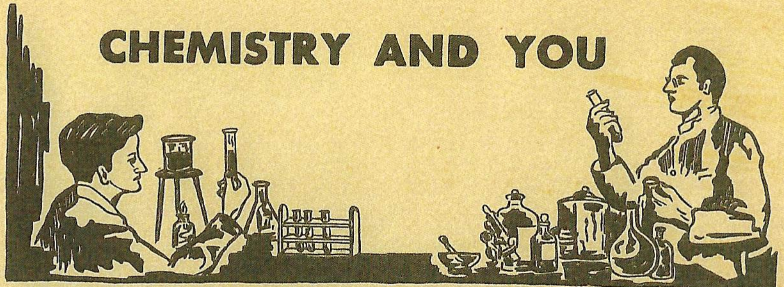
CHEMISTRY AND YOU



The entire universe about us, from its uppermost heights to its lowest depths, is built up of chemical elements and chemical compounds. Earth, sky and water are constantly passing through chemical changes. Deep in the ground, coal and oil are being formed from the remains of prehistoric forests and swamp vegetation. Precious metals and ores are being smelted under the heat and pressure of millions of tons of earth and rocks. On the surface of the earth, air and water are constantly producing chemical changes in everything they touch. All nature is merely a series of wonderful chemical reactions; plants, forests, birds, animals and people are all complex chemical engines.

Chemistry is more closely interwoven with the industries of the world than any other science, and the country which leads in chemical industries will ultimately be the richest and most powerful. It will have fewer waste materials, it will have the best manufactured articles, its food will be the most nourishing and the cheapest, and it will possess the secrets of the most powerful explosives, the hardest steels and the mightiest engines.

During the past score of years, the field of chemistry has become vastly enlarged. To illustrate how our basic industries, and we ourselves, rely on chemistry, there follows a partial list of everyday products and industries which depend directly or indirectly on chemistry for their present state of development:

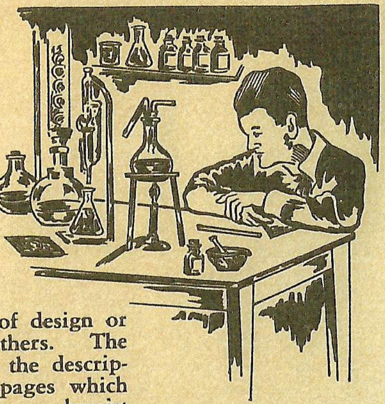
Foods
Construction
Fuels
Transportation
Medicine
Electronics
Inks and Dyes
Metallurgy
Plastics and Rubber
Adhesives
Engineering
Glass and Porcelain
Leather
Photography
Paints and Varnishes
Paper
Petroleum
Refrigeration and Air Conditioning
Soaps and Detergents
Textiles
Atomic Energy

Today, regardless of the profession a man follows, he is greatly handicapped without a knowledge of chemistry. The manufacturer, the farmer, the tradesman, the professional man, the scientist, all have constant need of chemical knowledge. In the home, the housewife is greatly handicapped if she does not have some understanding of the chemistry of the foods which she prepares and of the materials which she uses daily in her work. Mysterious chemical reactions even play an important part in enabling magicians to perform their tricks. In Chemcraft the many interesting branches of chemistry have been combined into a series of fascinating experiments which will furnish amusement for young chemists during many enjoyable and profitable hours.

The quantities of chemicals furnished with this outfit are sufficient to allow each experiment to be performed a number of times. An additional supply of any chemical or extra apparatus can be obtained by sending directly to the manufacturer. See the price list at the end of this book.

THE PORTER CHEMICAL COMPANY
Hagerstown, Maryland

HOW TO BE A CHEMIST



For a workman to be successful in any craft, there are two fundamental requirements which he must fulfill. In the first place, he must be able to use the tools of his trade with the proper skill, and in the second place, he must be able to understand and apply the principles of design or at least to follow the directions of others. The following important instructions and the descriptions of chemical apparatus on the pages which follow are designed to give the amateur chemist fundamental rules and information which will be the basis of developing skill in his craft.

Always read an experiment entirely through before starting its actual performance. By following this rule, many mistakes will be avoided.

Be sure to spread a thick layer of newspapers, oil cloth or other protecting material over the table on which you are experimenting so that liquids that may be spilled will not injure the surface.

All bottles and vials containing chemicals should be kept tightly stoppered as some chemicals become moist or gradually lose their strength when exposed to the air.

To pour a liquid from a bottle or vial, hold a glass rod against the bottom of the lip as you tip the bottle to pour from it. The liquid will then run down the rod and it is easy to control the quantity which comes from the bottle.

When using a flame, always set your candle or your alcohol lamp on a flat metal pan, such as a pie pan. This will avoid the danger of starting a fire should the candle or alcohol lamp be upset.

When you heat a liquid in a test tube, never point the mouth of it at yourself or at anyone nearby. The liquid may boil suddenly and spurt out of the tube.

When boiling a liquid in a test tube, hold the tube in a slanting position and shake or jiggle the tube slightly while heating. This will make the liquid boil much more evenly, but be careful not to slop the contents out of the tube.

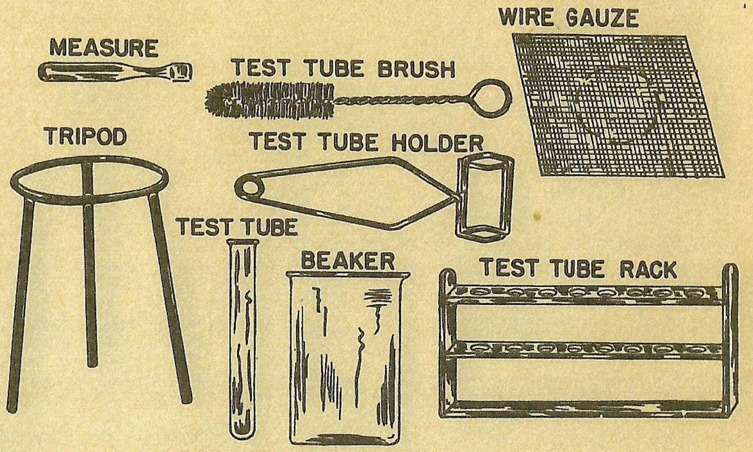
Always remove a test tube from the flame before smelling at the mouth of the tube to determine the odor of a gas.

When heating a solid in a dry test tube, warm the tube gradually. This can be done by slowly passing the test tube back and forth over the flame.

When heating a test tube over a candle, hold the bottom of the tube just above the tip of the flame. This will avoid the deposit of soot on the tube.

Cleanliness is of great importance in chemistry. Always keep your test tubes and other containers clean and wipe all of your equipment with a clean, dry cloth before putting it away.

The Use of Chemical Apparatus



THE USE OF CHEMICAL APPARATUS

This section describes the use of many pieces of chemical apparatus. Some of this apparatus may not be included in your set and the fact that it is described in this section does not mean that it is necessary for the experiments given in this manual.

MEASURE

"One measure of a chemical" means the quantity of a substance which can be held on the end of the small measure.

TEST TUBES

Test tubes are made of hard, thin glass, especially annealed to withstand heating. Liquids may be boiled and solids heated to high temperatures in test tubes, but care should be taken never to wet a dry, hot test tube, as this is almost certain to break it. When you have been heating a solid in your test tube, allow it to cool before washing it or adding a liquid.

TEST TUBE RACK

A test tube rack will be found very useful for holding your test tubes while experimenting.

TEST TUBE BRUSH

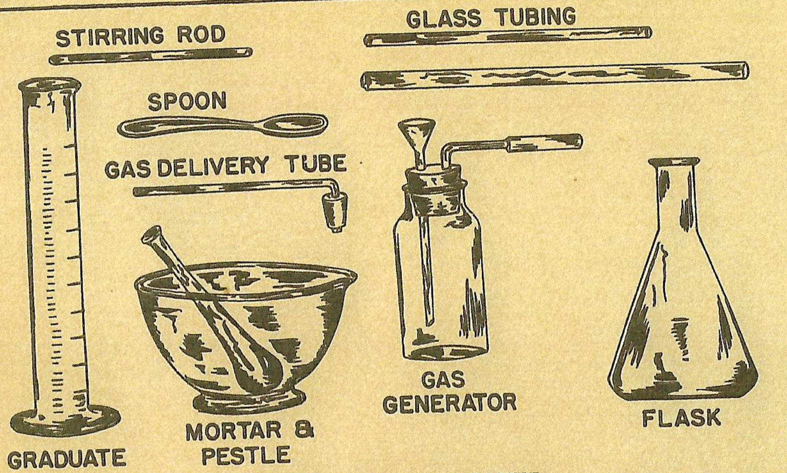
By filling a dirty tube half full of water and inserting the test tube brush, most substances can be easily removed. When test tubes are very dirty it is best to use hot water and rub the test tube brush on a cake of soap before inserting it.

TEST TUBE HOLDER

When heating substances in a test tube it sometimes becomes too hot to hold with your fingers.

BEAKER

A beaker can be used to hold hot liquids without danger of breakage. Liquids can also be heated in the beaker when supported on a wire gauze and tripod.



TRIPOD AND WIRE GAUZE

A tripod is used as a stand or support upon which you can set a vessel which you wish to heat. A wire gauze fits across the top of the tripod, making a convenient and safe stand for heating flat-bottomed vessels.

STIRRING ROD

Stirring with the glass rod often helps to dissolve a solid in a liquid. Always clean the glass rod by washing it in running water after taking it out of one liquid and before putting it into another.

GLASS TUBING

Glass tubing is used in setting up apparatus and in experiments on glass blowing.

SPOON

The spoon holds the equivalent of one teaspoonful, and is used for measuring liquids or solids. It is very convenient for heating compounds over the flame of a candle or alcohol lamp.

GAS DELIVERY TUBE

A gas delivery tube is used for the purpose of conducting a gas which is being generated in a test tube into any desired vessel.

MORTAR AND PESTLE

A mortar and pestle are used to grind or mix various substances.

GRADUATE

A graduate is used by the chemist for measuring amounts of liquids. The graduations etched in the glass wall of the vessel show you the quantity of liquid contained in the graduate.

FLASK

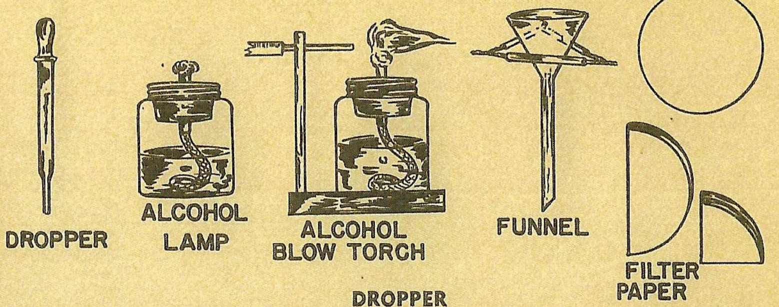
A flask will be found useful for heating liquids when the quantity is too large for a test tube. A flask is recommended when it is necessary to boil a liquid. Boiling liquids will not spurt out of a flask as easily as from a test tube. A flask should not be filled much over half full.

GAS GENERATOR

This piece of laboratory apparatus is assembled as shown in the above illustration. The funnel can be removed and used for filtering.

RUBBER TUBING

Short pieces of rubber or plastic tubing (2 inches) make excellent connectors for joining pieces of glass tubing when setting up chemical apparatus, a water bottle, or a siphon.



DROPPER

For transferring a measured number of drops of a liquid, a dropper is most convenient. Clean carefully by rinsing in clear water when you are through using it.

ALCOHOL LAMP

This is the cleanest and most convenient source of heat for most chemical experiments. Exercise great care in wiping up any alcohol that has been spilled before lighting your lamp. Do not fill the alcohol lamp near an open flame. Keep your alcohol bottle tightly stoppered and put it away as soon as you have finished filling your lamp.

ALCOHOL LAMP BLOW TORCH

With the Chemcraft alcohol lamp blow torch, it is possible to obtain quickly the high temperatures necessary in glass blowing and certain other chemical experiments. Set the alcohol lamp in the base block and insert the upright support, attaching the tube as shown in the diagram. Always be sure that the air jet tip (smallest hole) does not quite touch the flame. Adjust the air jet tip to the proper height so that when you blow into the tube, the flame is directed the greatest distance in a horizontal direction.

Blow into the tube with your cheeks puffed out, breathing through your nose. In this way you can maintain a steady stream of air from the reservoir in your cheeks. By placing the Chemcraft alcohol lamp blow torch on a suitable support, both hands are free for holding and manipulating the glass tubing or other work that is being heated in the flame.

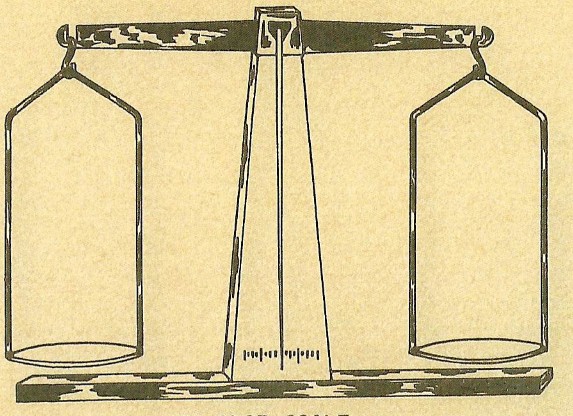
The blow torch can also be used with a candle flame where desired, by holding a candle so that the flame is in proper position with the air jet tip of the blow torch.

FUNNEL AND FILTER PAPER

It is often necessary to separate a solid from a liquid and this can usually best be done by filtering. A glass funnel and filter paper are generally used for this purpose, the filter paper being folded as shown in the diagram above. If necessary, a cone-shaped funnel can be improvised from a four-inch diameter circle of heavy wax paper. Fold this paper in half and then in half again. Cut off 1/8" at the bottom point, thus leaving a small hole at the end of the cone. Hold the wax paper in position by means of a paper clip. Now a piece of filter paper may be folded and placed into your wax paper funnel, or a small circular piece of newspaper can be used instead of filter paper. A wire triangle placed across a tumbler makes a fine support for a funnel.

BALANCE AND WEIGHTS

Balances of various types are used in every chemical laboratory. Some are used to weigh roughly large quantities of chemicals to be used in reactions; others weigh with extreme accuracy the small amounts of materials involved in chemical analysis.



BALANCE SCALE

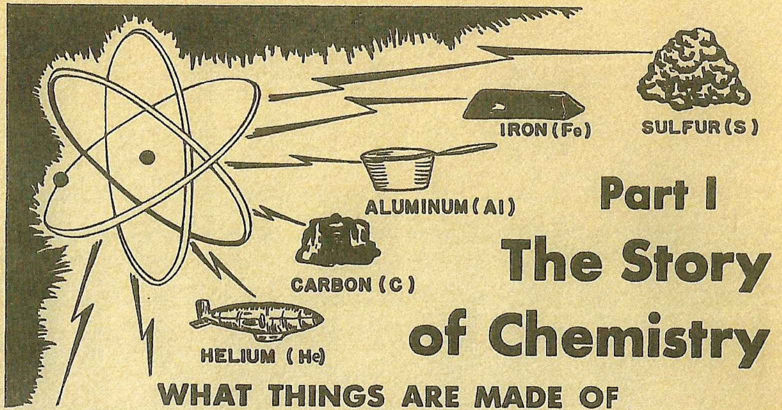
When weighing out solid chemicals, always place a piece of paper on each pan of the balance so that the chemicals will not injure the metal of the balance pan. Cut squares of smooth white paper, such as writing paper, so that the squares are of exactly the same size, and place one square on each pan of the balance. By having one on each pan, the papers counterpoise each other.

Now make sure that the beam of the balance is supported freely on its knife edges, and adjust the little metal rider on the beam so that the pointer swings an equal number of divisions on either side of the midpoint of the scale. When the pointer indicates that the instrument is properly adjusted, place the desired number of weights on the right-hand pan of the balance, and then add the chemical to be weighed to the paper on the left-hand pan until the swings of the pointer indicate that the weight of the chemical exactly equals the weights on the other pan. Remove the paper with the chemical on it from the left-hand pan.

In weighing liquids use small containers such as beakers. Place one on each pan of the balance and adjust the rider until they balance. If their weights differ considerably it may be necessary to add to one side or the other a few small weights such as coins in order to bring them into balance. When the pointer indicates that the pans are in balance, add the desired weights to the left-hand pan and pour the liquid to be weighed into the container on the right-hand pan until a balance is again reached.

In laboratory manuals and books of chemical recipes, the quantities of materials to be used are often given in terms of their weights. For example, in making the ordinary "50-50" solder, equal parts by weight of tin and lead must be melted together. You might weigh out, then, 10 grams of tin and 10 grams of lead if you wish to make 20 grams of solder.

If you need additional weights, it is easy to make them by cutting them out of sheet metal, such as aluminum or brass, and snipping them down until they just balance one or more of the weights which have been furnished. A new silver dime weighs almost exactly 2½ grams. A new nickel or 5 cent piece, weighs 5 grams. By adding several of these coins to your collection of weights, you will soon have quite a nice assortment. The gram is the unit of weight in the metric system and is used by chemists all over the world. There are about 28.35 grams in one avoirdupois ounce.



Part I

The Story of Chemistry

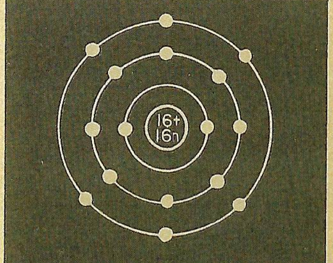
WHAT THINGS ARE MADE OF

Have you ever wondered of what things are made? When we look at a piece of iron or a pile of sulfur, they appear to be just that—iron or sulfur. Going more deeply, however, we find that all substances are constructed out of three kinds of units: (1) particles charged with negative electricity, called electrons, (2) particles charged with positive electricity, called protons, and (3) uncharged particles, called neutrons. All three of these units are extremely small, much too small even to be seen by the most powerful microscope.

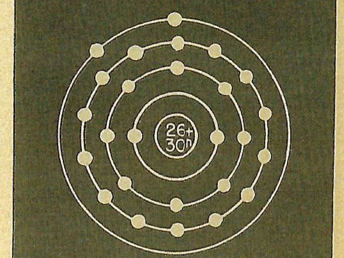
CHEMISTRY BEGINS WITH ATOMS

These three fundamental units—protons, electrons, and neutrons—may arrange themselves in various ways to form different kinds of atoms. An atom of iron, for example, is composed of 26 electrons, 26 protons and about 30 neutrons. An atom of sulfur is composed of 16 electrons, 16 protons and 16 neutrons. Chemists have in the past been principally concerned with the atoms themselves rather than with the smaller units of which they are constructed. This is because it has been necessary to apply enormous quantities of energy (using an "atom smasher") to change the structure of an atom itself, and this could only be done on a relatively small scale. The recent development of the atomic energy piles and the production of atomic bombs have greatly increased the importance of atomic chemistry, and already the conversion of one kind of atom into another is quite commonplace to the atomic chemist.

Any substance which contains only one kind of atom is called an element. Iron and sulfur are examples of elements. An element cannot be decomposed into any simpler substance by the usual chemical processes. There are about 98 different kinds of atoms, which means the same number of elements. Some of them are very familiar to you, but many are exceedingly rare and known only to scientists.



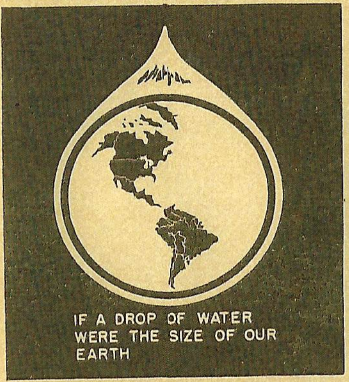
SULFUR ATOM



IRON ATOM

ATOMS COMBINE TO MAKE MOLECULES

Just as protons, electrons, and neutrons join to form atoms, atoms themselves may combine with other atoms. If atoms of two or more elements combine, the result is a molecule of a compound. For example, two hydrogen atoms joining with one oxygen atom form a molecule of water, represented by the formula H_2O. To get an idea of the size of a molecule of water, suppose you could magnify a single drop of water until it was the size of our earth. The individual molecules of water in the drop would then be about the size of basketballs, but the electrons in the atoms would still be too small to see without a microscope. Several hundred thousand kinds of molecules are known, some of them built up in a very complex way out of dozens of atoms. Sometimes two or more atoms of the same element combine, forming a molecule of the element. Thus the oxygen molecule consists of two oxygen atoms.


IF A DROP OF WATER WERE THE SIZE OF OUR EARTH

EXPERIMENT 1—Molecules Are Very Small.

Suppose we try an experiment that will give us some idea of how tiny a molecule is and how many of them are contained in a small space.

Obtain a fairly large clear glass container, such as a quart milk bottle, a large Mason jar, or a glass pitcher. Fill this container nearly to the top with water and let it stand quietly for 5 or 10 minutes, until the water has become perfectly still in the container. Now cut with scissors a small piece of your Dye Wafer (No. 8). This should be about one-quarter inch square. Drop this dye wafer on the surface of the water in the container and watch carefully for a few minutes.

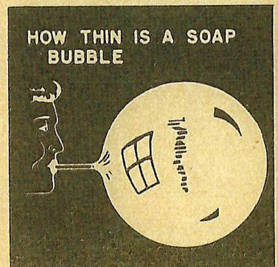
The dye wafer will float on the surface of the water and in a short time you will notice a streamer of color reaching downward through the clear water from the dye wafer. The dye is dissolving in the water and the colored molecules of the dye are diffusing through the liquid. If the water is thoroughly quiet the streamer of dye will remain concentrated until it reaches the bottom of the container, after which it will spread over the bottom and gradually diffuse upward.

After 15 or 20 minutes, shake the container slightly so that the molecules of dye are distributed evenly throughout the water. Notice that the water is tinted a decided color, showing that the small dye wafer which you used contained enough molecules to color a volume of water several hundred thousand times that of the dye.

Now empty half of the colored water from your container and refill it with clear water. Is the water still colored? Again empty half of the colored water and refill with clear water. How many times can you repeat this and still see a slight tint of color in the water? Remember that this color is caused by the colored molecules of dye and that all these molecules came from a small one-quarter inch square wafer of the dye. Does this give you an idea of how small a molecule must be?

EXPERIMENT 2—Molecules in a Soap Bubble Film.

Put a little water in a small bowl and then place a piece of soap in the water. Move the soap about with your fingers for a few minutes until some of it dissolves and the water becomes quite soapy. Now dip a bubble pipe or soda straw, or even a piece of glass tubing into the soap solution and carefully blow a soap bubble. As the soap bubble becomes larger you will notice a play of iridescent colors on the delicate film. The appearance of these iridescent colors indicates that the thickness of the bubble is only a few wave lengths of light, or only a few hundred-thousandths of an inch thick. Within this infinitesimal thickness, however, are molecules of water and molecules of soap. Just how many, we do not know, but at least this gives us some idea of how very, very small molecules must be.


HOW THIN IS A SOAP BUBBLE

MOLECULES ARE ALWAYS IN MOTION

The molecules in a substance are always moving or jiggling about. When we say that one object is hotter than another, we mean simply that its molecules are moving more energetically. When anything is heated, energy is added to it, and its molecules are made to move faster.

The molecules of a hot stove are in more rapid motion than when the stove is cool. When we touch the hot stove, its molecules jostle those of our skin, setting them, too, into abnormally rapid motion and causing us pain.

EXPERIMENT 3—The Motion of Molecules.

Fill a glass about ¾ full of water and add a teaspoonful of salt. Stir the liquid for a moment and the salt will entirely disappear. We say that the salt has dissolved in the water, which means that the crystals of salt have entirely separated and the molecules of salt being colorless, the water remains clear and of course the molecules of salt are too small to see.

Now repeat this experiment, using a glass ¾ full of fresh water, but before adding the teaspoonful of salt wait until the water becomes perfectly quiet. Pour the teaspoonful of salt into the water without stirring and the salt will settle to the bottom and will be visible there as a solid. Set the glass aside in a quiet place and let it stand overnight. The next day when you examine the glass again, much if not all of the salt will have disappeared. It has gone into solution in the water. If you hold the glass up to the light and swirl it very gently, you can see that there is a layer of dense salt solution in the lower part of the glass. Thus, although much salt has dissolved, it is not yet uniformly distributed throughout the water. There is, however, some salt even at the top, as you can easily prove by tasting a little of the water from the top of the glass. If the glass is covered and set aside for several days, the dissolved salt will slowly become distributed throughout the liquid, even without being stirred. This was largely due to the movement of the molecules of water and salt. These molecules constantly jiggling around gradually diffused the salt molecules into all portions of the glass of water.

THE ATTRACTION BETWEEN MOLECULES HOLDS THINGS TOGETHER

Molecules have an attraction for each other, that is, they tend to hang together. The soap bubble experiment which you performed is a good example of molecular attraction. Some kinds of molecules attract each other much more strongly than others. A brick, for example, is quite hard to break apart because the molecules of which it is composed attract each other strongly. So strongly do they hold on to each other that we can make them jiggle back and forth violently by heating the brick to a high temperature without forcing them to let go.

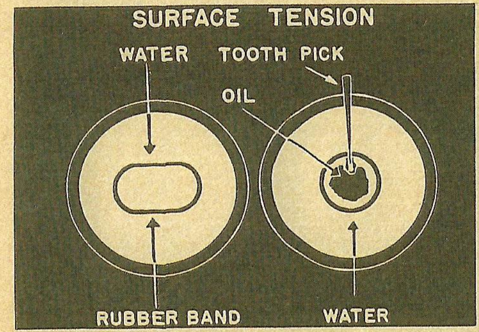
The best place to observe molecular attraction is in liquids, as all behave more or less as if their surfaces were covered with an elastic skin or film. This force, known as surface tension, is of course due to the attraction between the molecules of the liquid. It is surface tension or molecular attraction which causes a drop of liquid to assume a more or less spherical form as it falls through the air. The movement of soil moisture and of sap in plants and the concentration of ores by flotation all depend upon surface tension. Let us demonstrate this force between the molecules of a liquid with some simple experiments.

EXPERIMENT 4—The Floating Needle Shows the Attraction between Molecules.

Partly fill a saucer or small bowl with clean water. Obtain a small sewing needle, make sure that your fingers are perfectly dry, and carefully lay the needle gently upon the surface of the water. It will float there, held up by the elastic surface film due to the attraction between molecules.

EXPERIMENT 5—Demonstrating Surface Tension with a Rubber Band.

Carefully float a very light, thin rubber band on the surface of a dish of clean water that is free from any film of oil. Make sure that the band touches the water evenly all the way around. Dip a wire or toothpick into some vegetable oil, such as olive oil or salad oil, and then touch the water with the oily wire or toothpick inside of the rubber band. It may be necessary to repeat this operation several times, but you will soon notice that the band snaps out, tending to form a circle instead of its usual oval shape. This is because the oil lowers the surface tension or molecular attraction of the water inside the band, and the greater tension outside of the band then pulls on it from all directions.



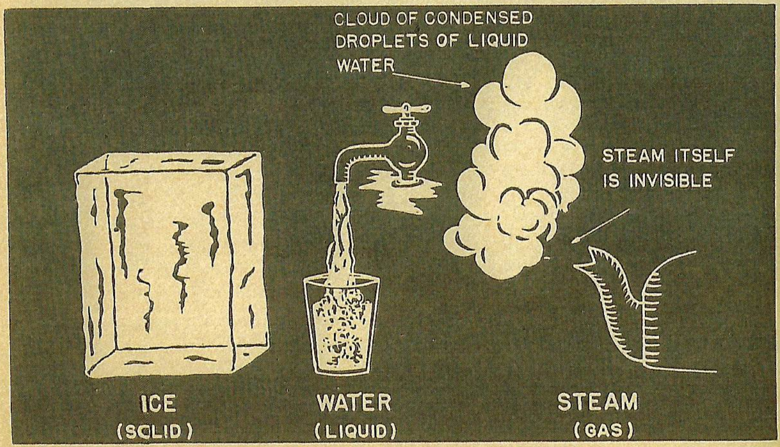
Now apply oil also to the water outside of the band. This will lower the surface tension outside so that the band springs back to its original shape.

The Difference between Solids, Liquids, and Gases

The real difference between the three states of matter - solid, liquid, and gas - is in the freedom of the molecules to move about. In solids the molecules are held in fairly definite positions, packed closely together, but their heat energy causes them to vibrate somewhat or jiggle back and forth. In liquids the attraction of the molecules for each other is not great enough to hold them firmly in place. They still stay close together, but the forces acting between them are so weak that they can slide over one another. Hence liquids can flow with ease and lack the strength and rigidity of solids. The molecules of a gas attract each other scarcely at all and hence remain far apart. Gases are, therefore, very light and have no strength or form at all.

EXPERIMENT 6—Changing a Solid to a Liquid and a Liquid to a Gas.

We have seen that when anything is heated its molecules move faster. Suppose we heat a piece of ice. Its molecules jiggle around harder and harder, until finally they jiggle right out of their places and slide around past each other. Thus the ice melts and becomes liquid. If heating the water is continued, making its molecules move still faster, every now and then one of them will be bumped so hard by its neighbor that it will sail right out through the surface and fly away among the molecules of gas which make up the air. This is why a liquid evaporates - because molecules occasionally are thrust out of it and escape. The hotter the liquid, the faster the molecules move, and the more chance they will have of escaping from the surface. That is why we warm things when we want to dry them quickly.


CLOUD OF CONDENSED DROPLETS OF LIQUID WATER
STEAM ITSELF IS INVISIBLE
ICE (SOLID)
WATER (LIQUID)
STEAM (GAS)

If water is heated strongly it will at last boil. This means that the molecules move so fast that they not only jump out of the surface very easily, but they even push each other aside in the body of the liquid and make hollow spaces - bubbles of steam, which rise through the liquid and escape. In steam, the molecules of water are whizzing about so fast that they have no chance to stick together, and hence a cloud of steam will not form a drop like liquid water, but spreads out and expands in all directions. This is why steam is so powerful for driving engines. The trillions of tiny molecules bump against the piston and push it in and out, moving trains and heavy machinery.

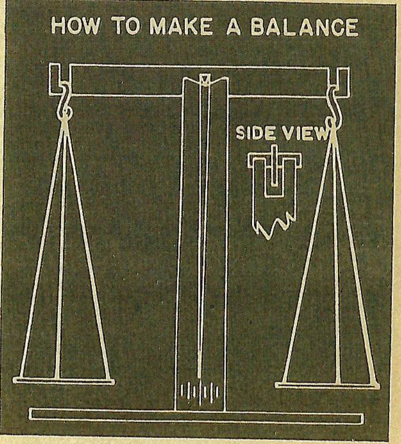
HOW CHEMISTS MEASURE MATTER

In all scientific work it is very necessary to have simple and accurate systems of measurement. In this country and in Great Britain, the English system is used for everyday measurement. This, as you know, consists of the inch, foot, and yard for linear measurement; the grain, ounce and pound for weight; and the pint, quart, and gallon for volume. In all other countries and for scientific work everywhere, the metric system is used. It is much simpler than the English system and adapts itself to scientific as well as to everyday measurements. The metric system is described and illustrated on Chemcraft Charts Nos. 2, 3, and 4.

Experiment 7—How to Make a Chemical Balance.

A chemical balance consists essentially of a base, a pedestal or upright, a beam which is pivoted at the center of the pedestal, and the pans which are suspended from each end of the beam. An indicator may be added if desired, but this is not essential.

A good size for the base is about 7″ long by 3″ wide by either ½″ or ¾″ thick. The pedestal or upright should be made out of a straight piece of wood either ¾″ square or 1″ square. This can be either 6″ or 7″ long and should be notched at the top so that the beam can be pivoted and will swing clear without touching the sides of the notch. The beam can be made of a piece of thin wood about as long as the pedestal is high, about ½″ wide, and not over ⅛ or $\frac{1}{16}$″ thick. If preferred, the beam can be made of a light, straight piece of metal in which case it will not have to be over $\frac{1}{16}$″ thick. The scale pans consist of round discs of metal or other hard material, either 1½″ or 2″ in diameter. Three strings of equal length and spaced equally around the outside diameter of the scale pans reach to the hooks which are suspended from each end of the beam.



The pivot on which the beam swings at the center can be a small, smooth iron nail. If it is desired to use a pointer, then the pivot should be triangular in shape so that the pointer can be forced onto the end of the pivot which extends to the front of the balance. In this case, small metal V's should be inserted in each side of the pedestal for the pivot to balance on. The pan hooks from the beam should be in line with the point of suspension of the beam itself, or if anything, a little below. If they are above, the beam will not swing properly, since the equilibrium at the center will be unstable.

A rider, which consists of a small piece of bent metal or even a small bent nail, is placed on the beam at the proper point to make the balance come to rest with the beam exactly level.

EXPERIMENT 8—How to Make Metric Weights.

A set of metric weights of fair accuracy can be made by using coins as standards. A new silver dime weighs almost exactly 2½ grams. Therefore, if you cut 5 equal size discs of aluminum which will exactly counterbalance a dime, each one of these discs will equal ½ gram. You can then cut larger discs which exactly counterbalance two of your ½ gram weights and these discs in turn will be 1 grams. A heavier piece of metal exactly equal to five 1 gram weights would be 5 grams, etc.

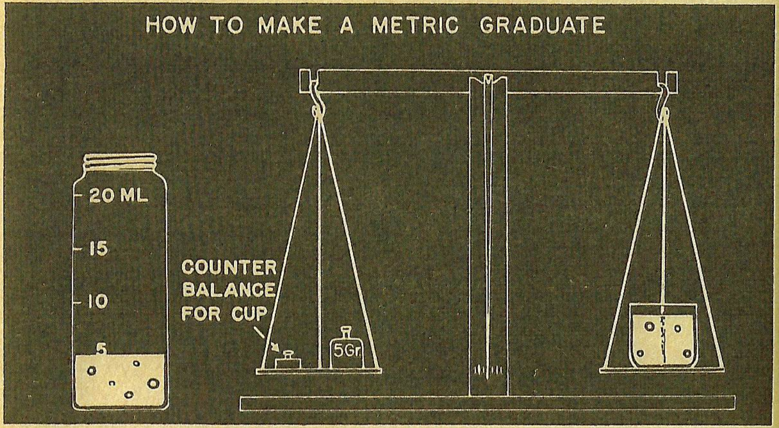
A nickel, or five-cent piece, weighs 5 grams. A copper penny weighs about 3.1 grams. A silver half dollar weighs about 12½ grams.

With your chemical balance and a set of metric weights, you will be able to weigh reasonably accurate quantities of chemicals. The desired number of weights, of course, are placed on one pan of the balance and the chemical added to the other pan until the balance comes to rest with the beam in a horizontal position.

EXPERIMENT 9—How to Make a Metric Graduate.

For this purpose you can use an empty shell vial, a tall thin olive bottle, or even a test tube.

Since 1 gram of water equals 1 cubic centimeter. or 1 milliliter, you can weigh out a certain known quantity of water and then using this volume as a standard, you can mark the graduations for your graduate with fair accuracy, using a sharp wax pencil to mark on the glass, or paste a strip of gummed paper on the glass. Use a small light cup on one side of your balance and enough weight on the other side to exactly balance the cup. The material which you use for counterbalancing does not have to be accurate weights, but can be any coins, or small bits of metal which will exactly balance the empty cup. Now place a 5 gram weight on the side of the balance away from the cup and pour enough water into the cup to again bring the balance to rest at exactly the right position. You now have exactly 5 grams or 5 milliliters of water in your cup. By carefully transferring this water to your graduate, you can place your 5 ml. mark accurately at the bottom of the curve of the water in the graduate. Two 5 ml. portions of water will then give you 10 mls. Two 10 ml. portions of water would give you 20 mls., etc., so that your graduate can be extended to any capacity which you wish.



EXPERIMENT 10—How to Make a Metric Scale.

A meter is approximately 39-⅜". Obtain a straight, thin piece of wood and cut it exactly 39-⅜" long. Now if you divide your stick in 10 equal parts you will have decimeters. This can thus be done by dividing the length exactly in half and then, with dividers, divide each half into exactly 5 equal parts. Now divide each decimeter in turn into 10 parts. These will be centimeters and you will then have a meter scale divided into decimeters and centimeters.

HOW CHEMISTS MEASURE TEMPERATURE

On a cold winter day you say, "Brrr, it's cold. The thermometer must be near zero." Or perhaps it's summer time, mid-afternoon in August with the sun shining bright, as you wipe the perspiration from your brow

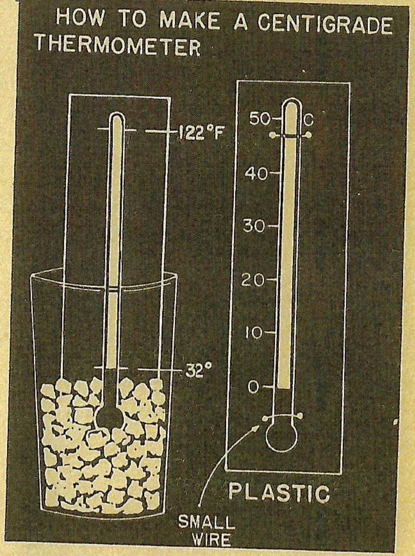
you say, "My, it must be 100° today." In both cases you are speaking of temperature in terms of the Fahrenheit scale on which the freezing point of water is 32 degrees and the boiling point of water is 212 degrees. Scientists, however, use a simpler scale for measuring temperature in which the freezing point of water is zero and the boiling point of water is 100 degrees. Chemcraft Chart No. 5 compares the Centigrade and Fahrenheit scales and shows you how to convert from one to the other.

EXPERIMENT 11—How to Make a Centigrade Thermometer.

If you have an old thermometer from which you can remove the tube, you can easily make a Centigrade thermometer, the kind which is used almost entirely in chemical work. Try to obtain a piece of plastic from which you can cut a strip about ½" wide and a little longer than the thermometer tube which you have. Drill a small hole for the hook at the top of the thermometer tube to fit in and grind out a space for the bulb near the bottom of the plastic strip. Now fasten your thermometer tube on the plastic back strip by drilling 2 sets of tiny holes and fastening the tube to the plastic by means of fine wire.

To graduate your thermometer, fill a tumbler with crushed ice, and add just enough pure water to fill the spaces between the chunks of ice. Set the thermometer so that the bulb is completely covered by the ice and let it stay in this position for at least five minutes. Now, accurately mark by means of a small scratch the exact point where the fluid within the tube stands. This will be 0 degrees Centigrade.

Now place your thermometer in a pan of water so that the bulb is again immersed and gradually heat the water to 122° Fahrenheit, adjusting the water to this temperature by means of another thermometer. Again mark the exact position of the fluid by means of a scratch on the plastic. This mark will be 50 degrees Centigrade. Now, by dividing the distance between the two marks in half and then dividing each half into 5's by means of a divider, your marks will then represent 5 degrees. If you have room, you can divide the spaces between the marks further into single degrees, but your marks in this case will be rather close together.



SPECIFIC GRAVITY

Do you float or sink in water? If you float, your specific gravity is less than one; if you sink, your specific gravity is more than one. If you neither sink nor float but just balance under the surface, your specific gravity is exactly one. Oil floats on water, therefore, it is lighter than water and its specific gravity is less than one (about 0.90). Mercury metal, of course, is much heavier than water and sinks immediately; its specific gravity is about 13.6. Specific gravity is quite important to the chemist in many ways. Chemcraft Chart No. 6 illustrates specific gravity and tells you how it is used.

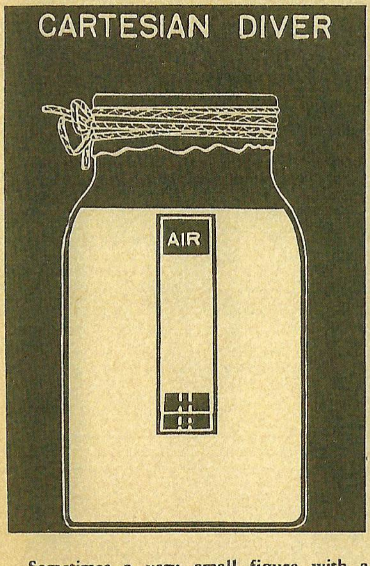
EXPERIMENT 12—Specific Gravity.

Test various household articles in a pail of water to see if you can find an object whose specific gravity is equal to one. If you can find such an object, it will neither sink nor float, but will hang suspended in your pail of water between the surface and the bottom of the pail. Some vegetables have a specific gravity very close to that of water. A fresh egg is just a little heavier than water, but a stale egg may float. Some kinds of fruit have a specific gravity very close to one. Some kinds of wood are quite light, but some of our heavier woods will actually sink in water. By driving nails into a small block of wood, it is possible to weight it so that it will have a specific gravity of exactly 1.

EXPERIMENT 13—Solution of Salt in Water Is Heavier than Pure Water.

If you have gone swimming in an ocean or in Great Salt Lake, you probably have noticed that you float better than you do in fresh water. This is because the dissolved salt increases the weight of the water. To illustrate this principle, fill two wide mouth fruit jars with water and in one of them, dissolve 3 or 4 teaspoonfuls of common table salt. Now, place an egg in the jar containing the pure water. If the egg is fresh, it will sink to the bottom of the jar. Place the same egg in the jar containing the salt solution and if you added enough salt, the egg will float on the surface. It is possible to add just enough salt so that the egg will remain suspended between the bottom of the jar and the surface of the solution.

EXPERIMENT 14—The Cartesian Driver.



Obtain a one or two quart wide mouth fruit jar. Fill the jar nearly full of water. Obtain a small vial not over 3 inches long and ½ inch in diameter. Make a hole through the cork, fill the vial about ¾ full of water and insert the cork. Now hold your finger over the hole in the cork, turn the vial upside down and lower it into your jar of water. As soon as the cork of the vial is under water, the air space in the top of the vial becomes sealed and will not change. You now want the vial to just barely float in the jar of water, that is, the bottom of the vial where the air is trapped should just barely touch the surface of the water in your jar. If the vial floats too high so that the bottom sticks above the surface of the water, remove it and add a little bit more water. If it sinks in the big jar, remove it and take out a little bit of water. Finally get it exactly balanced so that the bottom end of the vial barely touches the surface of the water in the jar.

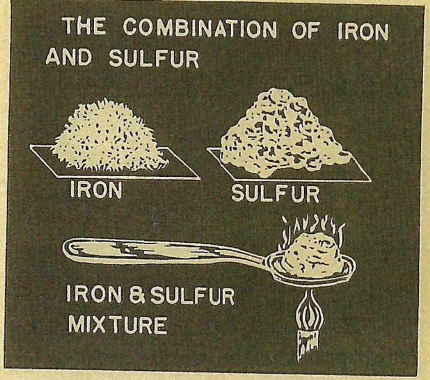
Obtain a square of rubber, such as a piece of inner tube, a little larger than the top of the jar. Tie the rubber securely over the mouth of the jar by wrapping a piece of string around several times and pulling it tight. Now if you press slightly on the piece of rubber, you compress the air in the top of the jar and this pressure is transmitted through the water through the hole in the cork of the small vial and slightly compresses the air that is trapped in the vial. When the air in the upper part of the small vial is compressed, it occupies less space, thus decreasing the buoyancy or increasing the specific gravity of the vial causing it to sink in the jar. If your vial is balanced exactly right, you will find that you can cause it to rise or sink in the jar at will by controlling the pressure which you exert upon the rubber covering the mouth of the jar.

Sometimes a very small figure with a small opening at the bottom and an air space inside is weighted so that it will just float in a jar of water. Then by means of the arrangement described above, these figures can be made to rise and fall at the will of the experimenter. This figure has been called the Cartesian Diver.

THE ELEMENTS

As we have learned, the elements are the fundamental substances with which the chemist builds. Ninety-eight of them are known at present, and these are listed on Chemcraft Chart No. 1. These elements can be combined with each other in many different ways, so the number of possible chemical compounds is enormous.

EXPERIMENT 15—Two Elements Combine to Form a Compound.


THE COMBINATION OF IRON AND SULFUR
IRON
SULFUR
IRON & SULFUR MIXTURE

Mix throughly on a sheet of paper 2 measures of Sulfur (No. 1) and 2 measures of iron filings or powdered iron metal. The powdered iron metal may consist of some steel wool cut by means of scissors into very small particles. By rolling the corners of the paper back and forth, you can mix the tiny pieces of iron thoroughly with the sulfur.

Separate the mixture into two equal parts and over one portion pass a magnet. You will find that you can pick the iron out of the sulfur in this way. The mixture can also be separated by washing the sulfur away with water, since it is lighter than the particles of iron.

Place the second portion of your iron and sulfur mixture in your spoon and heat it over a flame until it glows, or until the sulfur no longer burns. Keep your face away from the mixture while you are heating it. After the mass cools, examine it closely and see whether you can now separate the iron from the sulfur with a magnet or by washing it with water. A few stray particles of iron may have failed to combine, but most of the substance which you now have is neither iron nor sulfur, but a chemical compound of the two called iron sulfide or ferrous sulfide. You can represent the reaction which has occurred as follows:

$$\underset{\text{(iron)}}{\mathrm{Fe}} \quad + \quad \underset{\text{(sulfur)}}{\mathrm{S}} \quad \rightarrow \quad \underset{\text{(iron sulfide)}}{\mathrm{FeS}}$$

Each atom of iron becomes firmly tied up with an atom of sulfur and in order to again separate them, you would have to use another chemical reaction to reverse the effect of the first one.

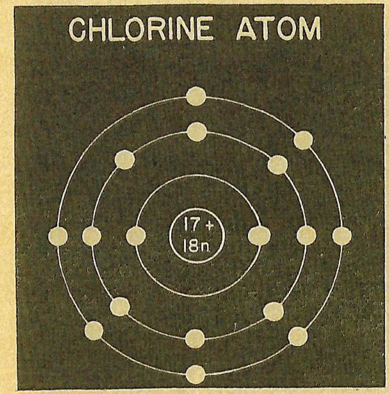
EXPERIMENT 16—A Compound Is Broken Down into Elements.

In the preceding experiment you made two elements combine to form a compound. Now let us see if we can break a compound down into its elements. Sugar is a compound. Suppose we heat 2 measures of sugar in a clean dry test tube and see what happens. Keep your face at a little distance and watch carefully the changes which take place in the sugar. Soon you will notice drops of water forming around the upper part of the test tube, therefore, sugar must contain water. Then the sugar becomes darker and finally only a charred substance remains in the test tube. This is the element carbon. Our experiment has shown that sugar is composed of water (hydrogen and oxygen) and carbon.

CHEMICAL SHORTHAND

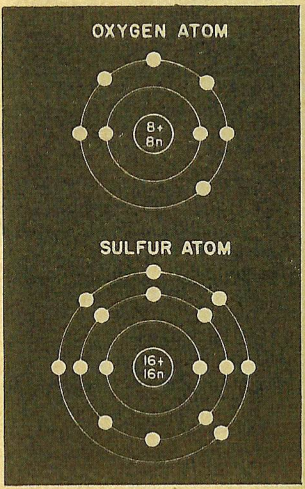
If you will look at the list of elements, you will notice three columns after the names of the elements. The first column gives the symbol of each element. These are abbreviations and are derived from the first letter or the first and a subsequent letter of either the English or Latin name of the element. The chemical name for common table salt is sodium chloride. The chemist would write NaCl, for natrium, the Latin for sodium, and Cl for chlorine.

HOW ATOMS ARE CONSTRUCTED



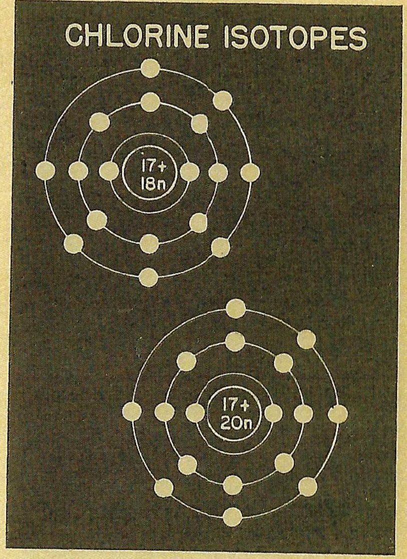
In order to understand how elements combine to form compounds and how compounds react with each other, we must return to the atom and see how it is constructed. An atom consists of a nucleus, or heavy center portion, around which are gathered very light electrons. The nucleus is made up of positively charged protons and uncharged neutrons. If we could enlarge an atom ten billion times, it would be about a yard across. Even with this tremendous magnification, the nucleus would be only a tiny pinpoint, and the individual protons and neutrons could probably not be seen. Small as it is, this tiny nucleus contains almost all the weight of the atom. Each of the electrons is about the size of the nucleus but weighs practically nothing. There are as many electrons outside the nucleus as there are protons in the nucleus, and they are believed to revolve about the nucleus very much as the planets of our solar system revolve around the sun. An atom is mostly empty space, somewhat like a swarm of mosquitoes, since the electrons are relatively far apart.

SOLVING THE MYSTERY OF ATOMIC WEIGHTS



In Chemcraft Chart No. 1, the List of Elements, you will find a column headed "Atomic Weights". Of course, an atom is entirely too small to weigh directly, but nevertheless chemists have been able to work out the relative weights of atoms of all the elements. To serve as a basis of comparison, the atom of oxygen was assumed to have a weight of 16 units, and the weights of the other atoms were compared with it by analyzing large numbers of compounds to find out the ratios by weight in which each of the elements combines with oxygen. Then, making allowance for the numbers of the different atoms in the compounds, it was possible to find the relation between the weight of each kind of atom and that of the oxygen atom. For example, we find by experiment that the compound sulfur dioxide contains equal weights of sulfur and oxygen. Now we know that in this compound there are only half as many atoms of sulfur as oxygen, and hence each sulfur atom must be twice as heavy as an oxygen atom if all together they are to weigh as much as double their number of oxygen atoms. Thus, letting the weight of the oxygen atom be 16, the relative atomic weight of sulfur is 32. Atomic weights are important because from them we can compute the relative weights of elements which enter into reactions.

WHAT ATOMIC NUMBERS MEAN

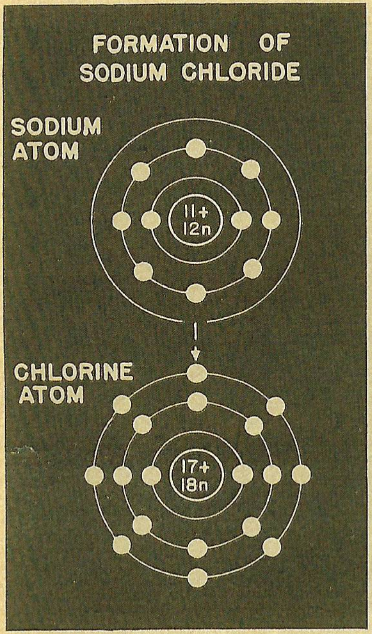


The column headed "Atomic Numbers" in Chemcraft Chart No. 1 tells us the number of protons in the nucleus of each kind of atom. These numbers are deduced by studying the x-rays emitted when each element in turn is made to serve as the target of an x-ray tube. Since the protons each carry one unit positive charge, their number tells us the positive charge of the atomic nucleus, and the number of electrons surrounding the nucleus must be exactly equal to this in a normal atom. Consider an atom of chromium. Its atomic number tells us there are 24 protons in its nucleus and hence 24 electrons around its nucleus. But its atomic weight is 52. Each proton has a weight of about one unit, and the electrons weigh almost nothing. Therefore, we know that the nucleus of chromium must contain enough neutrons (each with a weight of one unit) to build the weight up to 52. We picture, then, the chromium nucleus as composed of 24 protons and 28 neutrons. Try the same reasoning on the oxygen atom and see whether you come out with 8 protons and 8 neutrons. Many of the atomic weights are not even close to whole numbers. This simply means that not all the atoms of these elements have the same weight. Ordinary chlorine, for example, is a mixture of atoms having weights of 35 and 37. The atomic number of both kinds is 17, one contains in its nucleus 17 protons and 18 neutrons, and the other contains 17 protons and 20 neutrons. These two kinds of chlorine atoms are called isotopes. Some elements have as many as a dozen isotopes.

HOW ATOMS FORM IONS

The attraction of the positive nucleus for the negative electrons surrounding it binds the atom together. Certain atoms, however, may tend to lose one or more of their outermost electrons. Thus the hydrogen, sodium, and potassium atoms tend to lose one electron each; the calcium, magnesium, and copper atoms tend to part with two electrons; and the aluminum atom may give up three electrons. On the other hand, certain other atoms are so constructed that their tendency is toward gaining additional electrons rather than losing some which they already possess. Thus, the chlorine, bromine, and iodine atoms tend to gain one electron apiece, while the oxygen and sulfur atoms may gain two. Of course an atom which has gained or lost electrons is no longer a normal, neutral atom. An atom which has more than its normal number of negative electrons is negatively charged and is called a negative ion. One which has lost electrons is left with a positive charge and is called a positive ion.

HOW ATOMS CAN COMBINE



FORMATION OF SODIUM CHLORIDE

Suppose we bring together an electron-losing atom, such as sodium, and an electron-gaining atom, such as chlorine. An electron will then leave the sodium atom and go over to the chlorine atom. The sodium atom, Na, becomes a sodium ion, Na^+, and the chlorine atom, Cl, becomes a chloride ion, Cl^-. These ions, having opposite electrical charges, attract each other and tend to cling together, forming units in a solid crystal of sodium chloride, common salt, NaCl. It would, of course, be more accurate to write its formula Na^+Cl^- but we rarely bother to do so. A positively charged ion is called a cation and a negatively charged ion is called an anion.

When a substance is dissolved in water, it can react much more readily with other substances because the molecules of which the compounds or elements are composed are free to move about in the solution and combine with each other.

EXPERIMENT 17— To Show That Water Aids Chemical Reaction.

Mix thoroughly on a sheet of paper 1 measure of Strontium Chloride (No. 10) and 1 measure of Aluminum Sulfate (No. 12). Notice that there is no change in these chemicals and you can see no sign of any chemical reaction.

Now dissolve 1 measure of Aluminum Sulfate (No. 12) in a test tube ¼ full of water and dissolve 1 measure of Strontium Chloride (No. 10) in another test tube ¼ full of water. Pour these solutions together and you will immediately see a heavy white precipitate showing that the chemicals have reacted and have formed a new compound. This is because the ions which made up the crystals of strontium chloride and aluminum sulfate were separated from each other when the compounds were dissolved in water. They were thus free to enter into a new combination, forming the insoluble white compound, strontium sulfate.

EXPERIMENT 18—Another Chemical Reaction Depending on Ionization.

Mix 1 measure of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21) on a piece of paper and see if you can see any signs of a chemical reaction. Now pour the mixture of dry chemicals into a test tube half full of water and note that a chemical reaction immediately takes place, due to the formation of new compounds. Here again the ions have separated when the compounds are dissolved, and are then able to enter into a new combination.

WHAT IS AN ELECTROLYTE?

When a solid salt, for example sodium chloride, which is composed of sodium ions and chloride ions, is dissolved in water, the water molecules pry apart the ions, overcoming their attraction for each other. Thus the ions are free to float away in the liquid, and the solid salt crystal dissolves. A salt

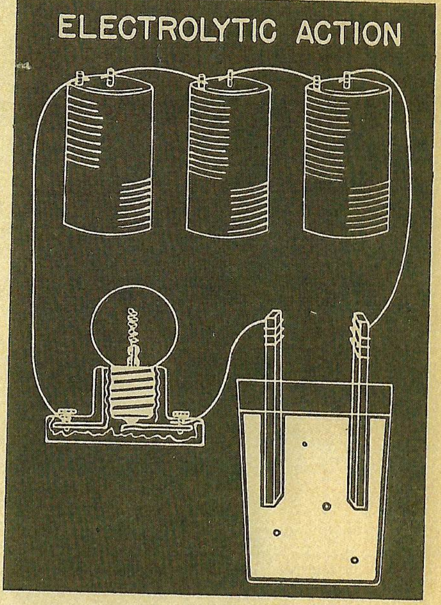
solution, therefore, contains no molecules of sodium chloride but only sodium and chloride ions, although for the sake of simplicity we referred to them in an earlier experiment as molecules. Such a solution, since it contains electrically charged ions, is a conductor of electricity and if wires leading from the poles of a battery are dipped into it, a current flows, carried through the solution by the ions.

Many other compounds resemble sodium chloride in general structure. The typical salts, such as sodium bromide, potassium nitrate, sodium phosphate, copper sulfate, barium chloride, etc., are composed of ions, as are bases, such as sodium hydroxide and calcium hydroxide. Acids in the pure state usually are not made up of ions, but as soon as they are dissolved in water a reaction occurs which produces many or few ions, depending upon whether the acid is strong or weak. Compounds like those we have been discussing are called electrolytes, since their solutions contain ions and hence conduct electricity.

EXPERIMENT 19—An Electrolyte Will Conduct Electricity.

Connect 3 dry cell batteries in series and hook a 3 volt flashlight bulb to one of the wires coming from the dry cells. Fasten a small zinc or copper plate, or a carbon rod, to the free end of each wire, as shown in the accompanying drawing. Now, fill a tumbler or a jelly glass about half full of water and put your plates or carbon rods in the glass so that they are partly immersed in the water. Move them back and forth, varying the distance between them and notice that your flashlight bulb will not light unless the plates or rods actually touch each other.

Now dissolve 1 or 2 teaspoonfuls of common table salt, sodium chloride, in the water and again introduce the plates or rods which are connected with your batteries. In this case, you will find the flashlight bulb will light even though the plates do not touch each other, showing that electric current is going through the solution.



ELECTROLYTIC ACTION

ANOTHER WAY IN WHICH ATOMS CAN COMBINE

Sometimes atoms combine with each other in a different way. Instead of one atom giving up completely an electron to another, two atoms may share a pair of electrons between them, one from each atom. Since both atoms are trying to hang on to the same pair of electrons, they stick together, forming a molecule of the compound. Some examples of this type of combination are water, H_2O; carbon dioxide, CO_2; alcohol, C_2H_6O; glycerine, $C_3H_8O_3$; and sugar, $C_{12}H_{22}O_{11}$. Such compounds are not composed of ions, and provided no reaction occurs when they dissolve, their solutions contain no ions and hence do not conduct electricity.

EXPERIMENT 20—Showing that Sugar is not Ionized.

Dissolve a teaspoonful of sugar in a glass half full of water, and, using the apparatus for the experiment on electrolytes, prove that a sugar solution will not conduct electric current.

CHEMICAL EQUATIONS

Now that we have learned how atoms form ions and combine with other ions to form molecules, this is a good place to find out how the chemist expresses these reactions by means of equations.

Before you write an equation for a reaction, you must know what substances react and what substances are formed by the reaction. Suppose you add sulfuric acid, H_2SO_4, to sodium bicarbonate (baking soda), $NaHCO_3$, and obtain as products carbon dioxide gas, CO_2, sodium sulfate, Na_2SO_4, and water. First write down the formulas for the reacting substances and the products:

$$NaHCO_3 + H_2SO_4 \rightarrow CO_2 + Na_2SO_4 + H_2O$$

Notice that in order to form a molecule of sodium sulfate you must put in two sodium atoms. To get two sodium atoms you need two molecules of sodium bicarbonate. Therefore, on the lefthand side of the equation write: $2\ NaHCO_3 + H_2SO_4$. Having taken two molecules of sodium bicarbonate, you know that you are going to get out two molecules of carbon dioxide on the righthand side, and if you count up the hydrogen and oxygen atoms, you will find that enough are left over to form two molecules of water. The balanced equation, then, is:

$$2\ NaHCO_3 + H_2SO_4 \rightarrow 2\ CO_2 + Na_2SO_4 + 2\ H_2O$$

No atom is ever lost in a chemical reaction. Atoms may be rearranged, re-shuffled, and re-dealt, but they are never destroyed and never created in ordinary chemical processes. This is the law of conservation of matter.

Chemical equations are simply concise ways of writing down what one actually finds out about a reaction by trying it. It must not be supposed that every nicely balanced equation that can be written will always mean something. It is easy to think up all sorts of reactions which look well on paper but do not actually work in the laboratory. The advanced branches of chemistry called thermodynamics and chemical kinetics tell us which reactions will actually react and which will not, but usually the easiest way to find out is to try the experiment and see.

METALS AND NON-METALS

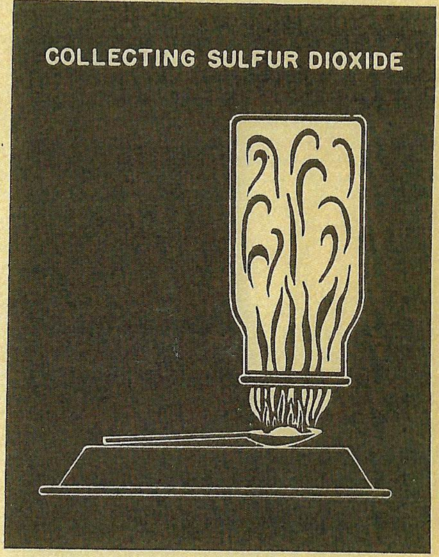
The elements can be more or less definitely divided into two classes — metals and non-metals. The metals are the elements whose atoms lose negative electrons, forming positive ions. Atoms of non-metallic elements tend rather to gain electrons, forming negative ions, or at least to acquire a share in some of the electrons of another atom. Some of the elements are hard to classify in this way since their atoms may either gain or lose electrons depending upon circumstances. Chemically, such elements have both metallic and non-metallic properties.

ACIDS AND BASES

In general, oxides (formed when metallic elements combine with oxygen) unite with water to form bases. Bases are compounds containing hydroxide ions, OH^-. When they are sufficiently soluble, their solutions turn red litmus paper blue. Most of the oxides of non-metallic elements, on the other hand, unite with water to form acids. The solutions of acids contain hydrogen ions, H^+, and turn blue litmus paper red.

EXPERIMENT 21—Making Sulfurous Acid.

Put 2 measures of Sulfur (No. 1) in your spoon and heat the spoon over a flame until the sulfur takes fire and burns. Now rest the spoon containing the burning sulfur on a metal plate and hold over it, mouth down, a one-pint milk bottle, so that the fumes from the burning sulfur will go into the bottle. After you have trapped some of the fumes, cover the mouth of the bottle with your hand and carefully pour about ½ inch of water into the bottle. Again close the mouth of the bottle tightly with your hand and shake for a moment so that the fumes within the bottle can be dissolved by the water. Now moisten a piece of blue Litmus Paper (No. 18) with the water from the milk bottle. It will turn red, showing that you have formed an acid. Sulfur, when burned in air, produces sulfur dioxide, which is an oxide of a non-metal. When this is dissolved in water, we obtain sulfurous acid.



$$\underset{\text{(sulfur dioxide)}}{SO_2} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(sulfurous acid)}}{H_2SO_3}$$

EXPERIMENT 22—Making Carbonic Acid.

Put 2 measures of baking soda in a test tube, add a few drops of vinegar and immediately place a gas delivery tube and stopper in the test tube. Fill another test tube half full of clear water and allow the gas being generated in the first tube to bubble through the water in the second tube. After the gas has bubbled through the clear water for 2 or 3 minutes, test the water in the second tube with a piece of blue Litmus Paper (No. 18). You will find that the blue litmus paper turns slightly pink, showing that you have a weak acid.

When you added vinegar to baking soda, carbon dioxide gas was given off and when carbon dioxide gas is dissolved in water it forms a weak acid known as carbonic acid.

$$\underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(carbonic acid)}}{H_2CO_3}$$

EXPERIMENT 23—Making a Base.

Put 2 measures of Calcium Oxide (No. 20) in a test tube and fill the tube about half full of water. Shake thoroughly to dissolve as much of the solid as possible. After the remaining solid has settled, test the clear liquid in the tube with a piece of red Litmus Paper (No. 18). The litmus paper will turn blue, showing that you have a base. You dissolved calcium oxide, which is an oxide of a metal, in water, a process which we have learned forms a base.

$$\underset{\text{(calcium oxide)}}{CaO} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(calcium hydroxide)}}{Ca(OH)_2}$$

WHAT WE MEAN BY THE pH OF A SOLUTION

It is frequently necessary to state just how acidic or basic a given solution is. Of course, this depends upon the quantity of hydrogen ions or of hydroxide ions in a unit volume (such as a liter) of the solution, but for many purposes it is inconvenient to state this quantity itself. Instead we derive from it by mathematical means a number called the pH of the solution, which tells at a glance how strongly acidic or basic the solution is. Chemcraft Chart No. 7 shows that the pH of pure water is 7, while the acidic solutions have lower pH values and the basic or alkaline ones have higher pH values.

The chart summarizes also the important uses of the pH system in chemistry, biology, medicine, engineering, and agriculture.

SALTS

Salts are compounds made up of positive and negative ions but containing neither hydrogen nor hydroxide ions. Common table salt, sodium chloride, is the most familiar example, but there are hundreds of others. A salt is formed whenever an acid and a base neutralize each other. For example, sodium chloride is formed when hydrochloric acid and sodium hydroxide react. Similarly, calcium sulfate is produced by the reaction of calcium hydroxide with sulfuric acid.

EXPERIMENT 24—Making a Salt.

Put 2 measures of baking soda in a test tube, add a few drops of vinegar and immediately place a gas delivery tube and stopper in the test tube. Fill another test tube half full of clear water and allow the gas being generated in the first tube to bubble through the water in the second tube. After the gas has bubbled through the clear water for 2 or 3 minutes, set this tube to one side.

$$\underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(carbonic acid)}}{H_2CO_3}$$

In another test tube place 1 measure of Calcium Oxide (No. 20) and fill the tube half full of water. Shake thoroughly to dissolve as much of the calcium oxide as possible and after the excess has settled to the bottom of the tube, pour the clear solution into another test tube.

$$\underset{\text{(calcium oxide)}}{CaO} + \underset{\text{(water)}}{H_2O} \rightarrow \underset{\text{(calcium hydroxide)}}{Ca(OH)_2}$$

You now have in the first test tube carbon dioxide gas dissolved in water. This is a non-metallic oxide dissolved in water, which we have learned forms an acid; in this case carbonic acid. In the second test tube you dissolved a metallic oxide in water which formed a base, calcium hydroxide. Now pour the contents of one of the tubes into the other and note that you obtain a white precipitate. You now have added an acid to a base, forming a salt.

$$\underset{\text{(carbonic acid)}}{H_2CO_3} + \underset{\text{(calcium hydroxide)}}{Ca(OH)_2} \rightarrow \underset{\text{(calcium carbonate)}}{CaCO_3} + \underset{\text{(water)}}{H_2O}$$

When a strong base, such as sodium hydroxide, combines with a weak acid, such as carbonic acid, the salt which is formed is basic or alkaline in reaction. On the other hand, when a strong acid, such as sulfuric acid, combines with a weak base, such as aluminum hydroxide, the resulting salt gives an acid reaction.

Another type of salt which may give an acid reaction in solution is the so-called "acid salt", formed when an acid is treated with too little base to neutralize it completely. An example of an acid salt is sodium bisulfate, formed when we add to sulfuric acid only half the amount of sodium hydroxide which would be required to entirely neutralize the acid.

EXPERIMENT 25—An Alkaline Salt.

Fill a test tube half full of water and add 1 measure of Sodium Carbonate (No. 21). Shake the tube for a moment to completely dissolve the solid. Now cut a small strip of red Litmus Paper (No. 18) and dip this into the test tube containing the sodium carbonate solution. The litmus paper will turn blue, showing that the sodium carbonate gives a basic or alkaline solution.

EXPERIMENT 26—An Acid Salt.

Again fill a test tube half full of water and add 1 measure of Sodium Bisulfate (No. 7). Cut a small piece of blue Litmus Paper (No. 18) and moisten this with the sodium bisulfate solution. The blue litmus paper will turn red, showing that sodium bisulfate gives an acid solution.

EXPERIMENT 27—A Neutral Salt.

Place 1 measure of sodium chloride (common table salt) in a test tube half full of water. Shake until the solid is entirely dissolved. Now test this solution with first red and then with blue Litmus Paper (No. 18). Does the color of either piece of litmus paper change? If not, this shows that sodium chloride is a neutral salt.

EXPERIMENT 28—Testing Vinegar.

Is vinegar an acid or a base? Test some vinegar with red and blue litmus paper, keeping in mind that if the blue litmus paper turns red, it is acid and that if red litmus paper turns blue, it is a base.

EXPERIMENT 29—Testing Lemon Juice.

Test a few drops of lemon juice in the same manner that you tested the vinegar. Does lemon juice contain an acid or a base?

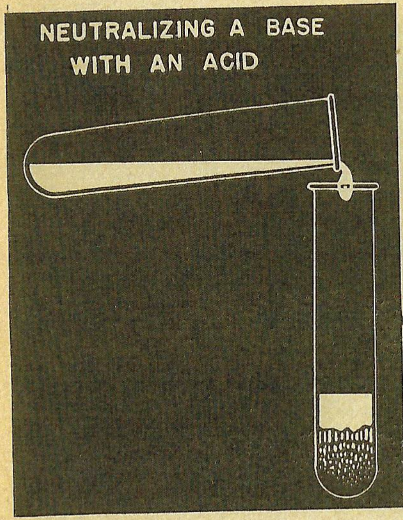
EXPERIMENT 30—Testing Ammonia Water.

Test a few drops of household ammonia with red litmus paper and blue litmus paper. Is ammonia an acid or a base?

EXPERIMENT 31—The Reaction of an Alkaline Salt with an Acid Salt.

Fill a test tube 1/4 full of water and add 1 measure of Sodium Carbonate (No. 4). Shake to dissolve the solid. In another test tube 1/4 full of water, dissolve 1 measure of Sodium Bisulfate (No. 7). Now pour the contents of one test tube into the other. The liquid fizzes up just like soda water. As we have learned, sodium carbonate gives a basic solution and sodium bisulfate an acid solution. When we brought them together, therefore, they tended to neutralize each other, forming sodium sulfate and liberating carbon dioxide.

EXPERIMENT 32—Neutralizing a Base with an Acid.


NEUTRALIZING A BASE WITH AN ACID

Place ½ measure of Calcium Oxide (No. 20) in a test tube half full of water and shake the tube well to dissolve as much of the calcium oxide as possible. As you have already learned, calcium oxide reacts with water to form calcium hydroxide, and so you now have a solution of calcium hydroxide. Set the test tube aside so that the undissolved solid can settle out.

In another test tube place 2 measures of Sodium Bisulfate (No. 7), fill it half full of water, and shake the tube until the solid is dissolved.

Now take a third clean test tube and pour into it some of the calcium hydroxide solution. The solid particles should have settled by this time, and if you are careful you can pour off just the clear part. Add to the clear solution 1 drop of Phenolphthalein Solution (No. 22) and notice that a clear red color appears, showing that the solution is alkaline, since the color change is a test for the presence of an alkali or base.

Next add some of the sodium bisulfate solution, a few drops at a time, to the calcium hydroxide solution; and if you work carefully, you will find that the red color becomes fainter until finally one more drop of sodium bisulfate solution, which is acid, will turn the phenolphthalein colorless. You now have neutralized the alkaline calcium hydroxide solution by adding acid, and the solution which you have in the test tube is now neither alkaline nor acid (at most, slightly acid). It is a neutral solution. The salt formed in this case is calcium sulfate.

INDICATORS

There are certain substances known as indicators, which have the property of turning one color in the presence of acids and another color in the presence of bases. They can thus be used to show whether any liquid is acidic, basic, or neutral.

There are many different kinds of indicators, and they are very interesting substances. The following experiments will introduce you to some of the common indicators.

EXPERIMENT 33—Litmus Paper.

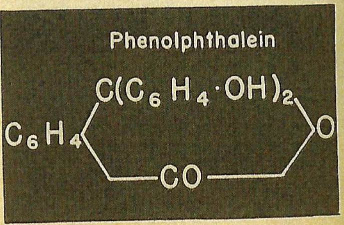
Put a piece of blue Litmus Paper (No. 18) into a glass half full of water. Add ½ measure of Sodium Bisulfate (No. 7) (or a few drops of vinegar) and stir. Next add 2 measures of Sodium Carbonate (No. 4) and stir again. Notice the color change.

EXPERIMENT 34—Phenolphthalein.

Put 1 or 2 drops of Phenolphthalein Solution (No. 22) into half a glass of water. Drop in one measure of Sodium Carbonate (No. 4), which is an alkali, and stir. Next drop in 2 or possibly 3 measures of Sodium Bisulfate (No. 7), which is an acid, or a few drops of vinegar (which contains acetic acid) and stir again.

Phenolphthalein is an indicator which is red in the presence of alkalies and colorless in the presence of acids. Some water supplies are alkaline and hence give the red color as soon as the phenolphthalein is added.

The molecule of phenolphthalein is a complicated structure built up of atoms of carbon, hydrogen, and oxygen joined together as shown in the diagram.



EXPERIMENT 35—Phenolphthalein Paper.

Dip a piece of soft, white paper into Phenolphthalein Solution (No. 22) and spread it out to dry. When it is dry, you have a piece of paper which you can use as an indicator merely by moistening it. Place ¼ measure of Sodium Carbonate (No. 4) on the paper and add a drop of water. Notice the red color which forms immediately.

Moisten a piece of the phenolphthalein test paper and hold it over the mouth of a bottle of household ammonia. It will turn red, showing that ammonia fumes are alkaline.

EXPERIMENT 36—Testing a Laxative Tablet for Phenolphthalein.

Obtain a stick of laxative chewing gum or a laxative tablet and break it up into small pieces. Put these into half a glass of water and stir for a minute or two to extract some of the phenolphthalein. Now add one measure of Sodium Carbonate (No. 4) and if the tablet contains phenolphthalein, the liquid will turn red.

Many of the most common flowers, fruits, and vegetables with which you are familiar contain coloring matter which belong to a class of indicators. Try some of the following experiments if you can obtain the necessary flowers, fruits, or vegetables.

EXPERIMENT 37—Hollyhock Indicator.

The purple hollyhock contains a coloring matter which makes a very good indicator. Break up several purple hollyhock flowers and place them in a tumbler or small dish. Now add to them about ½ glass of hot water and allow them to stand with occasional stirring until the liquid becomes a deep purple.

EXPERIMENT 38—Hollyhock Test Paper.

Moisten a piece of soft, white paper with the purple hollyhock liquid and when it is dry, use it to test for acids or alkalies. This paper is called "Mallow Paper" and it is frequently used as a test paper because it is quite sensitive to the presence of either acids or alkalies.

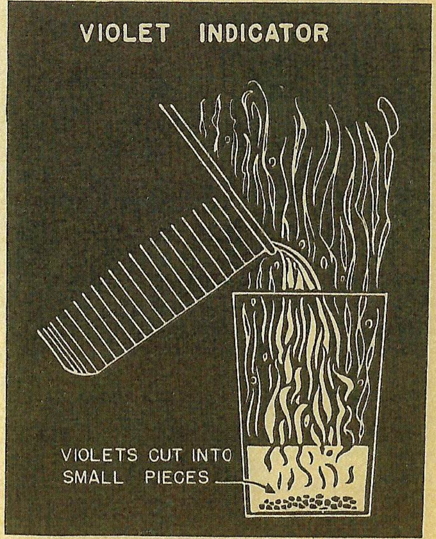
When you make these different test papers, it will be a good plan to label them so that you know what each one is later on.

EXPERIMENT 39—Violet Indicator.

Violets contain a coloring matter that they may be used as an indicator. Suppose you pick about two dozen of these flowers, remove the stems and cut the blossoms into tiny pieces. Now extract the coloring matter with hot water as you did in Experiment 37. When you obtain a good greenish-blue solution, test a small portion with some Sodium Carbonate (No. 4) dissolved in water, and then with some vinegar. Notice the sharp color changes. The coloring matter in violets is red in acids, and green in alkalies.

EXPERIMENT 40—Violet Test Paper.

To make violet indicator test paper, moisten a piece of soft, white paper with your violet liquid prepared in the preceding experiment. When your paper is dry, test it with an alkali (sodium carbonate) and an acid (vinegar). Here again you will notice that the color changes as it did in the experiment on a violet indicator.


VIOLET INDICATOR
VIOLETS CUT INTO SMALL PIECES

EXPERIMENT 41—Flag Indicator.

The blue flag or iris contains a coloring matter which may be used as an indicator. If you can obtain some blue flags, put them in a glass and extract the coloring matter with hot water. When a deep blue solution is obtained, test a small portion of it with a little sodium carbonate and then with a little sodium bisulfate or vinegar and notice the color change.

EXPERIMENT 42—Dahlia Indicator.

If you can obtain some purple dahlias, a very good indicator can be made from them. Break the flowers into small pieces and pour half a glass of hot water over them. When the solution has become a deep purple, test a small portion of it with ¼ measure of Sodium Carbonate (No. 4) and also with a little acid.

The coloring matter contained in purple dahlias is believed to be an anthocyanin, very closely related to the dye known as "Hoffman's Violet". The coloring matter from the dahlia is green with alkalies and red with acids.

EXPERIMENT 43—Dahlia Test Paper.

Moisten a piece of soft, white paper with a little of the purple dahlia prepared in the preceding experiment. When dry it should be a rich blue violet color. This paper can be used to test for acids or alkalies; being red with acids and green with alkalies.

EXPERIMENT 44—Cherry Indicator.

Obtain some cherries; the dark red sweet ones are best for this experiment. Place a number of cherries in a dish and crush the juice from them. Test a little of this juice with ¼ measure of Sodium Carbonate (No. 4) and then add 1 measure of Sodium Bisulfate (No. 7) and notice the color changes with alkalies and acids.

EXPERIMENT 45—Elderberry Indicator.

Elderberries contain a juice which can be used as an indicator. If you can obtain a few elderberries, crush them and strain the juice. Test this juice with 1 measure of Sodium Carbonate (No. 4) and then with 1 measure of Sodium Bisulfate (No 7). You will find that this indicator is greenish blue with alkalies and red with acids.

EXPERIMENT 46—Elderberry Test Paper.

If you can extract enough of the elderberry juice, put it in a test tube and moisten a piece of soft, white paper with it. When the paper is dry, test a small strip of it with alkalies and acids.

EXPERIMENT 47—Huckleberry Indicator.

If you can get some huckleberries during huckleberry season, put a few in a dish, add a little water and grind the mass until a rich liquid is obtained. Pour off a little of this liquid into a clean test tube and test with sodium carbonate, and then with sodium bisulfate or vinegar. Notice the different colors.

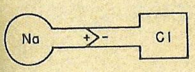
EXPERIMENT 48—Red Cabbage Indicator.

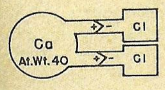
An indicator may be prepared from red cabbage leaves. If you can obtain a red cabbage, break up a few of the leaves into small pieces and soak them overnight in a tumbler half full of water. Stir this liquid until it becomes colored a deep purple. It may be necessary to put the cabbage leaves in hot water in order to extract enough of the color.

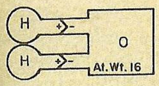
Now pour a little of the clear purple liquid into a test tube and add ¼ measure of Sodium Carbonate (No. 4). You will find that the solution has turned green. Now add 1 measure of Sodium Bisulfate (No. 7), and you will find that the green liquid turns red. This indicator is green in the presence of alkalies and red with acids.

COMBINING CAPACITY OR VALENCE

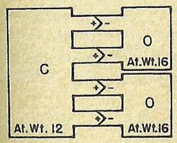
Atoms of some elements, as we have learned, can lose one or more negative electrons while atoms of other elements can gain various numbers of negative electrons. Perhaps there are some questions which should now be answered. Why should sodium chloride be NaCl but calcium chloride $CaCl_2$? If water is H_2O, why should carbon dioxide be CO_2 and not C_2O? Years of patient experimenting have shown that each element has a definite ability or capacity for combining with other elements. This combining capacity we call valence. A hydrogen atom has one unit of combining capacity, that is, its valence is one. Since one atom of chlorine can combine with one, and only one, atom of hydrogen in forming the compound HCl, the valence of chlorine must also be one. Furthermore, one atom of sodium combines with one atom of chlorine to form NaCl, and so the sodium must likewise have a valence of one.

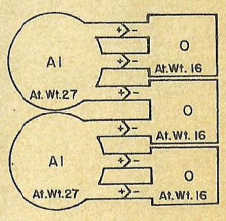




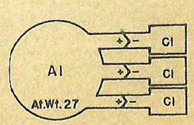


It must not be overlooked, however, that when a sodium atom combines with a chlorine atom, it gives up an electron to the chlorine and hence becomes positively charged, while the chlorine receives a unit of negative electrical charge. We say, therefore, that sodium has a valence of plus one, and that chlorine has a valence of minus one. Hydrogen, like sodium, has ordinarily a valence of plus one, since it, too, tends to give up an electron. Now what about calcium? It has been shown that it can combine with two chlorine atoms by giving up two electrons to them. Its valence then is plus two. Since aluminum can form the compound aluminum chloride, $AlCl_3$, its valence must be plus three. In water there are two atoms of hydrogen for every one of oxygen, and so oxygen must have twice the combining capacity of hydrogen, since each atom of it is able to hold two of hydrogen. The valence of oxygen is thus minus two. How would you write the formula of sodium oxide? Would it be Na_2O? What about calcium oxide? Since calcium has a valence of plus two and oxygen a valence of minus two, each can hold just one atom of the other, forming the compound CaO. Aluminum oxide may be a little harder. If you take two aluminum atoms, each having a valence of plus three, you will have a total positive valence of six. Now how many oxygen atoms must you have in order to get a total negative valence of six? It will take three atoms of oxygen to furnish enough negative combining power to hold the two aluminum atoms. Therefore, the formula of aluminum oxide is Al_2O_3. This compound, colored by impurities, forms the gem minerals, sapphire, ruby, and corundum.





The number of electrons lost by the atoms of some metals may depend upon circumstances. An atom of iron, for example, may lose either two electrons or three electrons. If it loses two and if each is taken up by a chlorine atom, it forms ferrous chloride, $FeCl_2$. If it gives up three, one to each of three



chlorine atoms, ferric chloride, $FeCl_3$, will be formed. Iron, then, is an element with two valences, plus two and plus three. When its valence is two, its compound with oxygen is FeO, ferrous oxide. When its valence is three, it can hold still more oxygen, forming Fe_2O_3, ferric oxide. A number of the elements have two or more valences.

THE DIFFERENCE BETWEEN PHYSICAL CHANGE AND CHEMICAL CHANGE

When water freezes, it is still water. When it evaporates and forms vapor, it is still built up of the same atoms of hydrogen and oxygen. When sugar dissolves in water, it is still sugar. The atoms of carbon, hydrogen, and oxygen remain in the same arrangement in the molecules, and if the water is removed by evaporation the residue is solid sugar. On the other hand, when a wax candle burns, it is no longer wax, and it would be a hard job to "unburn" the candle and get back the original wax. When a piece of iron rusts, new molecules are formed from atoms of iron and atoms of oxygen. Changes like this, which alter the composition of compounds by adding, removing, or rearranging atoms, are called chemical changes. Changes of the first kind, which affect only the physical state of a substance without altering the structure of its molecules, are called physical changes. In more advanced courses in chemistry you will learn that it is occasionally very difficult to decide whether a certain change should be called chemical or physical, but it is of little importance to consider such cases here.

EXPERIMENT 49—A Physical Change.

Put a teaspoonful of sodium chloride, ordinary table salt, in a tumbler half full of water and stir until the salt is all dissolved. Now pour this solution into a small sauce pan and heat it on the stove until the water completely evaporates. When the salt becomes nearly dry, there may be some popping, but this does no harm other than to cause some of the salt to fly around.

Examine carefully the salt that remains in the pan and satisfy yourself that it is the same as the salt with which you started. We learn, therefore, that merely dissolving a substance does not necessarily produce a chemical reaction, even though the physical form of the substances changes completely.

EXPERIMENT 50—A Chemical Change.

Put 2 measures of Azurite (No. 39) (copper carbonate) into a clean dry test tube. Heat this dry powder over your alcohol lamp flame for a few minutes and notice what occurs. First, water is given off as you can see from the moisture formed in the upper part of the test tube. Next, the copper carbonate will gradually become black as it is converted into black copper oxide. In this case, the copper carbonate has been changed chemically, being broken down into copper oxide and carbon dioxide, and you no longer have the original substance with which you started.

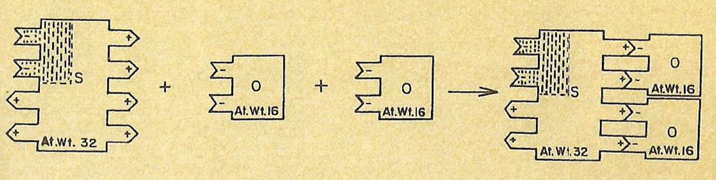
$$\underset{\text{(copper carbonate)}}{CuCO_3} + \text{heat} \rightarrow \underset{\text{(copper oxide)}}{CuO} + \underset{\text{(carbon dioxide)}}{CO_2\uparrow}$$

CHEMICAL REACTIONS — COMBINATION OF ELEMENTS

One of the most common chemical changes is the combination of one element with another to form a compound. If a small quantity of sulfur is lighted, it will burn with a bluish flame until the sulfur is all gone. The oxygen of the air has combined with the sulfur forming sulfur dioxide, which is a gas:

$$S + O_2 \rightarrow SO_2\uparrow$$

Since an atom of sulfur in this case has a valence or combining capacity of 4, it requires 2 atoms of oxygen with a valence of 2 to unite with it and the sulfur dioxide molecule which is formed contains 1 atom of sulfur and 2 atoms of oxygen.

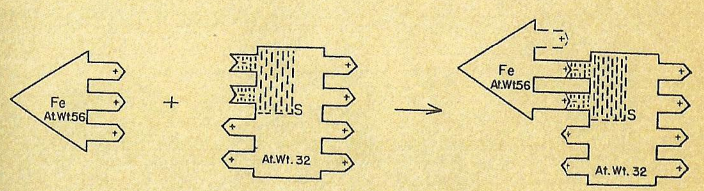


EXPERIMENT 51—The Combination of Iron and Sulfur.

Powdered iron metal, iron filings or steel wool cut up into very small particles will combine with sulfur when heated, forming iron sulfide. Mix together thoroughly on a sheet of paper 2 measures of finely divided iron and 2 measures of Sulfur (No. 1). Heat this mixture in the spoon until the sulfur no longer burns and a glow is noted in the mass. Keep your face away while you are heating. Examine the mass after it cools and you will find it is an entirely different substance from either iron or sulfur. If a small portion of the iron sulfide is put into the test tube with 1 measure of Sodium Bisulfate (No. 7) and a little water and the solution warmed, you will notice the odor of hydrogen sulfide, proving that you have iron sulfide present.

$$\underset{\text{(iron)}}{Fe} + \underset{\text{(sulfur)}}{S} \rightarrow \underset{\text{(iron sulfide)}}{FeS}$$

Notice that in this reaction sulfur has a valence of 2, while iron (ferrous) also has a valence of 2.

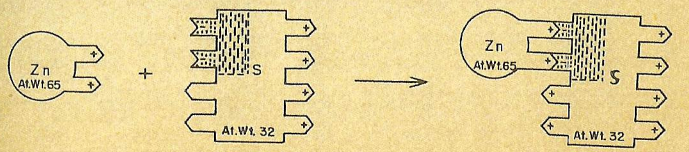


EXPERIMENT 52—Combination of Zinc and Sulfur.

Cut a few very fine slivers from a sheet of zinc. Mix these with an equal quantity of Sulfur (No. 1) and heat the mixture over a flame. The reaction in this case is more noticeable than the sulfur and iron and the product, of course, is zinc sulfide. This may be proven by adding a dilute acid and heating to obtain the odor of hydrogen sulfide as in the preceding experiment.

$$\underset{\text{(zinc)}}{Zn} + \underset{\text{(sulfur)}}{S} \rightarrow \underset{\text{(zinc sulfide)}}{ZnS}$$

Here, also, sulfur has a valence of 2.



EXPERIMENT 53—Combination of Zinc and Oxygen.

Cut a fine sliver of zinc and grip one end of it in your test tube holder. Now, holding the test tube holder in your hand, introduce the other end of the sliver of zinc into a flame. Notice that the zinc burns with a bluish flame or in other words, unites with the oxygen of the air, forming a white powdery substance, which is zinc oxide.

$$\underset{\text{(zinc)}}{2\ \mathrm{Zn}} + \underset{\text{(oxygen)}}{\mathrm{O_2}} \rightarrow \underset{\text{(zinc oxide)}}{2\ \mathrm{ZnO}}$$

EXPERIMENT 54—Combination of Sulfur and Oxygen.

When 1 or 2 measures of sulfur are put in the spoon and heated until the sulfur catches on fire and burns, the sulfur is combining with the oxygen of the air, forming sulfur dioxide, SO_2.

$$\underset{\text{(sulfur)}}{\mathrm{S}} + \underset{\text{(oxygen)}}{\mathrm{O_2}} \rightarrow \underset{\text{(sulfur dioxide)}}{\mathrm{SO_2}\uparrow}$$

CHEMICAL REACTIONS — DECOMPOSITION

We have seen above how 2 elements combine. Now we shall illustrate how a molecule can be split up or decomposed into its elements. By passing an electric current through water the hydrogen can be separated from the oxygen:

$$2\ \mathrm{H_2O} \rightarrow 2\ \mathrm{H_2} + \mathrm{O_2}$$

In this reaction O_2 is used to designate a molecule of oxygen gas because each molecule of the gas is made up of 2 atoms. It is, therefore, necessary to have 2 molecules of hydrogen, each of which consists of 2 atoms, in order to balance the equation.

EXPERIMENT 55—Decomposition of Sodium Thiosulfate.

Place 2 measures of Sodium Thiosulfate (No. 16) in a test tube. Use the test tube holder so as not to burn your fingers, and heat the bottom of the tube over a flame. Notice that the solid first becomes moist and steam is given off. This is because sodium thiosulfate contains water of hydration which is driven off when the substance is heated.

Continue heating for a few minutes longer and you will notice that a yellowish powder, sulfur, is being deposited on the upper part of the test tube. Remove the tube from the flame and smell the gas which is given off at the mouth of the tube. You will recognize an odor like rotten eggs due to the gas, hydrogen sulfide.

You learn from this experiment that by heating sodium thiosulfate, it is broken up into several different substances. First — water, which was driven off as steam. Second — sulfur, which was deposited around the upper part of the test tube. Third — hydrogen sulfide gas, which you could smell at the mouth of the tube. Fourth — a reddish mass in the bottom of the tube which consists of a mixture of sodium sulfate and sodium sulfide.

EXPERIMENT 56—The Decomposition of Sugar.

Place 2 measures of powdered or granulated sugar in a clean dry test tube. Heat the tube over a flame, keeping your face at a little distance, and watch carefully the changes which take place in the sugar. First — steam is given off, showing that the sugar contains water. Next the sugar gradually becomes darker and finally you have remaining in the test tube a black substance which is carbon, an element. Thus you have learned that sugar is composed of water (hydrogen and oxygen) and carbon.

$$\underset{\text{(sugar)}}{\mathrm{C_{12}H_{22}O_{11}}} + \text{heat} \rightarrow \underset{\text{(carbon)}}{12\ \mathrm{C}} + \underset{\text{(water)}}{11\ \mathrm{H_2O}}$$

EXPERIMENT 57—Decomposition of Calcium Carbonate.

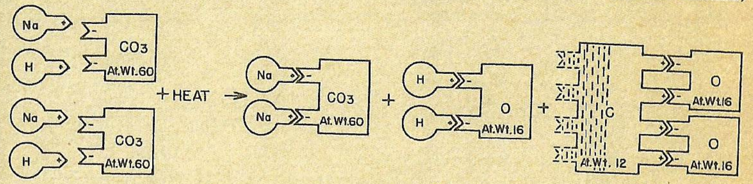
Calcium carbonate is the main constituent of limestone and oyster shells and can be decomposed by heat to form lime and carbon dioxide. Obtain a small piece of limestone or oyster shell and break it with a hammer so that you have a number of very small slivers. Pick out one of these slivers which has a very thin knife-like edge or point. Hold the sliver with a pair of pliers and heat the thin edge or point in your alcohol lamp flame for several minutes, keeping it red hot during this time. If you have a gas flame available, this experiment can be done more quickly with a hotter flame. After the sliver of limestone or oyster shell has cooled, examine carefully the edge which you heated and you will notice that this particular part has become white and chalky and that some white powder can be scraped off with a knife. The heat has driven off the carbon dioxide from the calcium carbonate, leaving calcium oxide, commonly known as lime.

$$\underset{\text{(calcium carbonate)}}{\mathrm{CaCO_3}} + \text{heat} \rightarrow \underset{\text{(calcium oxide)}}{\mathrm{CaO}} + \underset{\text{(carbon dioxide)}}{\mathrm{CO_2}\uparrow}$$

EXPERIMENT 58—The Decomposition of Sodium Bicarbonate.

Put 2 measures of sodium bicarbonate (baking soda) in a test tube and heat over a flame. You will notice that moisture is given off although the baking soda with which you started was perfectly dry. As you heat the sodium bicarbonate, light a match and thrust it into the mouth of the test tube. It will immediately go out, showing that carbon dioxide gas is being given off. After a few minutes a white powdery mass will remain in the test tube, which consists of sodium carbonate.

$$\underset{\text{(sodium bicarbonate)}}{2\ \mathrm{NaHCO_3}} + \text{heat} \rightarrow \underset{\text{(sodium carbonate)}}{\mathrm{Na_2CO_3}} + \underset{\text{(water)}}{\mathrm{H_2O}} + \underset{\text{(carbon dioxide)}}{\mathrm{CO_2}\uparrow}$$



To prove that you have decomposed sodium bicarbonate, dissolve 1 measure of baking soda in a clean test tube half full of water. Add 1 drop of Phenolphthalein Solution (No. 22) and notice that you get only a very faint pink color, showing that sodium bicarbonate is almost neutral. Now add a little water to the test tube containing the residue remaining after you heated the sodium bicarbonate. Make sure this test tube is cool before you add water. When this substance has dissolved, add 1 drop of phenolphthalein solution and note that you get an intense red color, showing that you have a strongly alkaline salt which corresponds with sodium carbonate.

CHEMICAL REACTIONS — DISPLACEMENT

When a metal is acted upon by an acid, the hydrogen of the acid is displaced by the metal and the hydrogen is liberated as a gas:

$$\mathrm{Zn} + \mathrm{H_2SO_4} \rightarrow \mathrm{ZnSO_4} + \mathrm{H_2}\uparrow$$

The metal in the above reaction is zinc which displaces the hydrogen of the sulfuric acid forming zinc sulfate and liberating hydrogen. Zinc has a valence or combining power of 2 and hydrogen has a valence or combining power of 1, therefore, 1 atom of zinc replaces 2 atoms of hydrogen in the sulfuric acid molecule H_2SO_4. This then becomes zinc sulfate, $ZnSO_4$, and the 2 atoms of hydrogen are liberated in the form of a hydrogen molecule.

EXPERIMENT 59—Displacement of Hydrogen by Zinc.

Dissolve 2 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water. Now cut several fine slivers of zinc metal, using an old pair of scissors or tin snips. Drop these slivers of zinc into the test tube containing sodium bisulfate solution and warm the solution over a flame. You will notice bubbles of gas beginning to come from the pieces of zinc. This is hydrogen being displaced from the sodium bisulfate while the zinc is going into solution in the form of zinc sulfate. After this action has been going on for a few moments, hold a lighted match at the mouth of the test tube. If the hydrogen gas is coming off fast enough, you will notice a slight pop as the hydrogen is ignited.

$$\underset{\text{(sodium bisulfate)}}{2\,NaHSO_4} + \underset{\text{(zinc)}}{Zn} \rightarrow \underset{\text{(sodium sulfate)}}{Na_2SO_4} + \underset{\text{(zinc sulfate)}}{ZnSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 60—Displacement of Hydrogen by Iron.

Dissolve 2 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water and add a few iron filings or very small pieces of steel wool. Warm this solution over a flame and you will notice bubbles of gas forming on the iron. In this case the action is not as vigorous as it was with the zinc because iron is a less active metal than zinc.

$$\underset{\text{(sodium bisulfate)}}{2\,NaHSO_4} + \underset{\text{(iron)}}{Fe} \rightarrow \underset{\text{(sodium sulfate)}}{Na_2SO_4} + \underset{\text{(iron sulfate)}}{FeSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 61—Displacement of Copper by Iron.

Put 1 measure of Azurite (No. 39) in a test tube and add 3 measures of Sodium Bisulfate (No. 7). Fill the test tube half full of water, close the mouth with your thumb and shake until the solids are entirely dissolved.

Azurite is a mineral containing a large percentage of copper carbonate. This substance while insoluble in water will dissolve readily in sodium bisulfate solution, forming a solution of copper sulfate (blue vitriol).

Now add 1 measure of powdered iron metal, or small pieces of steel wool, to the solution of copper sulfate which you have just made and shake the tube for a few minutes. The blue color of the copper sulfate solution will gradually disappear and the iron will become covered with a red deposit of copper.

Iron compounds, when dissolved in water, give a blue color with Sodium Ferrocyanide (No. 6). Add ½ measure of this substance to the test tube and shake it. This will prove that some of the iron is now dissolved in the liquid.

$$\underset{\text{(copper sulfate)}}{CuSO_4} + \underset{\text{(iron)}}{Fe} \rightarrow \underset{\text{(ferrous sulfate)}}{FeSO_4} + \underset{\text{(copper)}}{Cu}$$

From the above equation you can easily see what happened. The iron simply pushed the copper out of the copper sulfate, forming ferrous (iron) sulfate and setting free the copper in the metallic state.

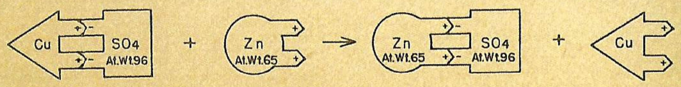
EXPERIMENT 62—Displacement of Copper by Zinc.

Prepare some copper sulfate solution by adding 3 measures of Sodium Bisulfate (No. 7) and 1 measure of Azurite (No. 39) to a test tube half full of water as in the preceding experiment.

Now add a small piece of zinc about the size of the end of the measure to the copper sulfate solution, close the mouth of the test tube with your thumb and shake it, then let stand for a few minutes. This time you will find that the zinc has become covered with a dark brown deposit of copper and the blue of the copper sulfate solution has disappeared.

$$\underset{\text{(copper sulfate)}}{CuSO_4} + \underset{\text{(zinc)}}{Zn} \rightarrow \underset{\text{(zinc sulfate)}}{ZnSO_4} + \underset{\text{(copper)}}{Cu}$$

This reaction, you see, is exactly like the preceding one except that it is zinc which displaces copper, forming, of course, zinc sulfate instead of ferrous sulfate.



CHEMICAL REACTIONS — DOUBLE DECOMPOSITION

Suppose some sodium carbonate (washing soda) is dissolved in water and added to a solution of calcium chloride. Immediately the liquid becomes milky with fine particles of a solid compound, calcium carbonate (precipitated chalk). If this solid is separated from the liquid by filtering, or by allowing the precipitate to settle, and the water is evaporated from the remaining clear liquid, crystals of sodium chloride (table salt) will be left in the vessel. Here is what happened, expressed in the form of a chemical equation:

$$CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2\,NaCl$$

The calcium and the sodium seemed to have exchanged partners during the reaction.

This equation we have just written is good enough for most purposes, but it does not tell quite the whole story, because compounds such as calcium chloride and sodium carbonate are made up of ions. Each calcium atom has a strong tendency to part with two of its electrons, and each chlorine atom has an equally strong tendency to pick up one electron. It is easy to see, then, how two chlorine atoms might each take one of the electrons which a calcium atom is anxious to get rid of. This forms two negatively charged chloride ions and one positively charged calcium ion. The positive and negative charges attract each other, and in a crystal of solid calcium chloride the ions hold each other tightly together. When the crystal is dissolved in water, the ions are permitted to separate, and so there are calcium and chloride ions wandering around through the liquid.

In the same way, when sodium carbonate is dissolved, the solution contains sodium ions, Na^+, and carbonate ions, CO_3^{--}. Notice that the carbon atom and three oxygen atoms of the carbonate group hang together in a single ion and go through the whole reaction without breaking apart.

What really happens in this experiment, then, is the mixing of sodium ions and carbonate ions with calcium ions and chloride ions. The calcium ions immediately combine with the carbonate ions to form the insoluble compound, calcium carbonate:

$$2\,Na^+ + CO_3^{--} + Ca^{++} + 2\,Cl^- \rightarrow CaCO_3 + 2\,Na^+ + 2\,Cl^-$$

The sodium and chloride ions simply remain in the solution, and when the water is evaporated they come together to form crystals of solid sodium chloride.

A great many of the common reactions which go on in solution are really reactions between ions, but frequently one does not bother to show the ions in the equations.

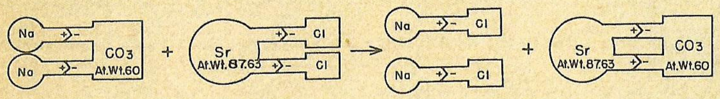
EXPERIMENT 63—Precipitation of Strontium Carbonate from Sodium Carbonate and Strontium Chloride.

Put 1 measure of Sodium Carbonate (No. 4) in a test tube. Fill the tube ¼ full of water and shake it until the solid is dissolved. In another test tube put 1 measure of Strontium Chloride (No. 10). Fill the tube ¼ full of water and shake the tube to dissolve the solid.

You now have a solution of two different salts. Pour the contents of one test tube into the other and notice that a white solid or precipitate is formed. The sodium carbonate and strontium chloride react forming sodium chloride and strontium carbonate. The sodium chloride was soluble in water and remained dissolved, but the strontium carbonate was insoluble in water and so it separated from the solution as a solid or precipitate.

$$Na_2CO_3 + SrCl_2 \rightarrow 2\ NaCl + SrCO_3$$

(sodium carbonate) (strontium chloride) (sodium chloride) (strontium carbonate)



EXPERIMENT 64—Precipitation of Strontium Sulfate from Aluminum Sulfate and Strontium Chloride.

Dissolve 1 measure of Aluminum Sulfate (No. 12) in a test tube ¼ full of water. In another test tube ¼ full of water dissolve 1 measure of Strontium Chloride (No. 10). You now have two colorless solutions. Pour the contents of one tube into the other and you immediately obtain a heavy white precipitate, showing that a chemical reaction has taken place. In this reaction you have formed strontium sulfate, which is insoluble in water and appears as a heavy white precipitate, while aluminum chloride remains in the solution.

$$Al_2(SO_4)_3 + 3\ SrCl_2 \rightarrow 3\ SrSO_4 + 2\ AlCl_3$$

(aluminum sulfate) (strontium chloride) (strontium sulfate) (aluminum chloride)

Part II

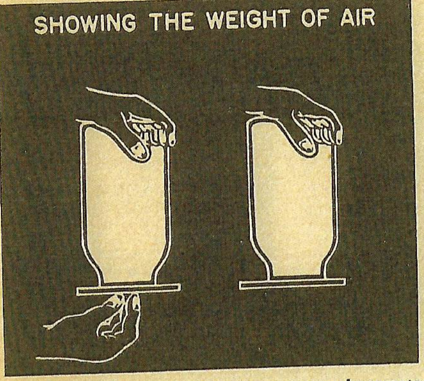
Our Chemical World

Our world and the whole great universe about us is built of chemical elements and compounds. On the surface of the earth, air and water are constantly producing chemical changes, while deep down in the ground, coal, oil, and ores of valuable metals are being formed under the heat and pressure of millions of tons of earth and rock. All life involves a series of wonderful chemical reactions; plants, forests, birds, animals and people are all complex chemical engines.

The Air Around Our Earth

Our earth is completely surrounded by a blanket or layer of gases. These gases are what we call air. Everyone realizes the importance of the air. It makes possible both animal and plant life; it keeps the sun from scorching us during the day and it reduces the earth's loss of heat into empty space during the night. This blanket of air is not very thick, as it extends upward from the earth's surface only about 50 miles. The weight of the air, however, over the whole surface of the earth is enormous. You may have heard that the atmosphere exerts a pressure of nearly 15 pounds per square inch at the surface of the earth. In other words, a column of air, one inch square and as high as the atmosphere extends, weighs almost 15 pounds. Thus the air directly above an acre of ground (an area about the size of the average baseball grounds) weighs approximately 90,000,000 pounds.

EXPERIMENT 65—Showing the Weight of Air.


SHOWING THE WEIGHT OF AIR

Obtain a milk bottle or drinking glass with a perfectly smooth edge. Working at a sink or wash basin, fill this container with water so that the water is even with, or stands just a tiny bit higher than, the rim of the glass or bottle. Now, cut a small piece of cardboard large enough to cover the mouth of the glass or bottle and carefully slide this across the top so that it forms perfect contact with the water inside of the container. Place the finger of one hand lightly on the cardboard and with the other hand turn the glass or bottle upside down. Now carefully remove your finger and if you have followed directions correctly, the water will not spill from the container. This is because the weight of the air presses against the cardboard strongly enough to prevent the water falling out of the glass or bottle.

The weight or pressure of the air is strong enough to support a column of mercury 30 inches high. This is the principle upon which the mercury barometers for measuring the air pressure work.

EXPERIMENT 66—The Weight of the Air Crushes a Metal Can.

If you can obtain a large metal can with a tight screw cap, you can illustrate the weight of the air very easily. The can must have a capacity of at least a gallon and of course must be tight, and also have a screw cap which fits well, or a tight-fitting wooden plug or rubber stopper. (A gallon size square oil can is ideal.)

Put enough water in the can to form a layer about an inch deep in the bottom. Now set the can on a stove and heat it until the water inside is boiling briskly and steam is coming out of the opening on top. At this point, turn off the heat or remove the can quickly from the burner and immediately screw the cap on tightly. Now, holding the can with a cloth so as not to burn yourself, hold it under a cold water faucet and allow the cold water to quickly cool the can. This condenses the steam inside the can which had largely replaced the air and as the steam is condensed, leaving a partial vacuum inside of the can, the weight of the air on the outside will cause the can to collapse or curve inward. If you use a large can, it may completely collapse, but a small can will probably show only slight dents in the side.

OXYGEN

What is this important mixture of gases that we call air? It consists principally of two gases, nitrogen and oxygen. About one-fifth of air is oxygen and four-fifths of air is nitrogen. There are several other gases present in small amounts, including argon, neon, krypton, and of course, carbon dioxide and water vapor.

It is the oxygen in the air which supports life. When we breathe, the oxygen of the air taken into the lungs is absorbed by the blood and carried to all parts of the body. The various tissues slowly unite with the oxygen, forming carbon dioxide. Heat is liberated by this reaction. The carbon dioxide is expelled in our breath. If the air should be deprived of oxygen, we could not live, but on the other hand, if there were no nitrogen in the air and we were forced to breathe pure oxygen, our tissues would be consumed so rapidly that we would soon expend our energy.

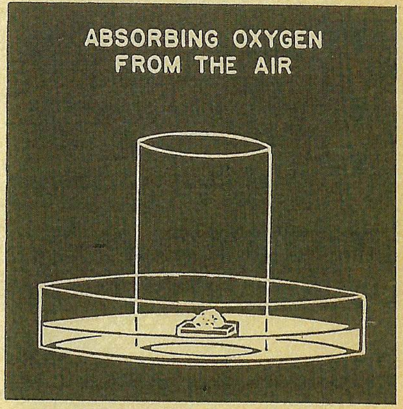
Plants absorb carbon dioxide from the air through their leaves, and combine it with water to form, by a series of complicated reactions, various compounds from which their tissues are very largely built up. In this process oxygen is liberated and returned to the air. Thus, plants and animals mutually assist in keeping the quantity of oxygen and of carbon dioxide in the air constant.

Of all the elements, oxygen is the most abundant and the most widely distributed. This element, in the free state, is a colorless, odorless gas. Air, as we have learned, is one-fifth free oxygen, and water, which we will soon take up, is eight-ninths combined oxygen. Common materials like sandstone, limestone, brick and mortar contain about 50% oxygen. It has been estimated that about one-half of our world is oxygen.

Oxygen is a very active element, combining readily with many other elements. These reactions are called oxidation. Burning or ordinary combustion is a process of oxidation. Thus, the burning of coal is mainly the oxidation of carbon.

It can readily be seen that the activity of oxygen affects our lives beneficially in many ways and yet, because of its abundance in the free state in the air and its high chemical activity, oxygen also creates grave problems. Uncontrolled combustion, such as fires that destroy our buildings and sweep our forests, and many of the disastrous explosions that wreck mines and buildings, sometimes causing great loss of life, are examples of the uncontrolled reactions of oxygen. The corrosion of metals, as, for example, the "rusting" of iron, really is a reaction of oxidation, or very slow combustion.

EXPERIMENT 67—Absorbing Oxygen from the Air.



Obtain a small flat piece of wood or cork which will fit loosely inside an ordinary tumbler. Fill a saucer or small dish with water, float the piece of wood on the water and on the wood, place a small wad of steel wool about the size of the end of your thumb. Make sure that it is steel wool (iron) and not some other metal such as copper. You can use 4 measures of powdered iron or iron filings if more convenient. Moisten the steel wool with a few drops of water. Now invert a tumbler over the floating wood in order to trap the air in the tumbler and then set the saucer and tumbler and floating wood in a quiet place for a few days.

After 2 or 3 days, when the steel wool has begun to rust, notice that the water is slowly rising inside of the glass, showing that the volume of the enclosed air is becoming less. This is because some of the oxygen in the air that was trapped inside of the glass has united with the iron of the steel wool, forming iron oxide or rust.

EXPERIMENT 68—Measuring the Amount of Oxygen in the Air.

You can easily determine the amount of oxygen in ordinary air by the following procedure:

Set up the apparatus and material that you used in the last experiment and before setting it aside, measure the distance from the surface of the water to the inside top of the glass. Record this distance and then set the glass aside as it will take some time for the iron to absorb all of the oxygen.

After several days, again measure the distance from the surface of the water to the top of the glass. To find the percentage of oxygen, subtract this value from the original reading, divide by the original reading and then multiply the result by 100. This should give you a value of approximately 20% for the amount of oxygen in air. The exact value cannot be determined because no corrections have been made for changes in temperature and pressure which take place during the experiment and which cause changes in the volume of the enclosed air.

EXPERIMENT 69—Removing Oxygen from Air with a Candle Flame.

Set a candle in the middle of an ordinary pan, first allowing a few drops of melted wax to fall on the pan to stick the candle tight. Next pour about 2 inches of water in the pan. Now light the candle and set over it an inverted quart milk bottle or fruit jar. The bottle or jar must be high enough so that the bottom will not come too close to the candle flame.

Notice that almost at once the candle flame grows dim and that after a few minutes it goes out, the oxygen of the air inside the bottle having been consumed. Notice also that after the flame has gone out the water begins to rise inside the bottle and finally stands at a considerably higher level in the bottle than in the pan. This is due to the removal of the oxygen from the air inside of the bottle, leaving a partial vacuum so that the air pressure on the outside pushes some of the water from the pan up into the bottle. In the flame, the oxygen of the air unites with the candle wax, which contains carbon, thus forming carbon dioxide, a gas which will dissolve in water, and steam which condenses into water.

EXPERIMENT 70—How Leaves Give Off Oxygen.

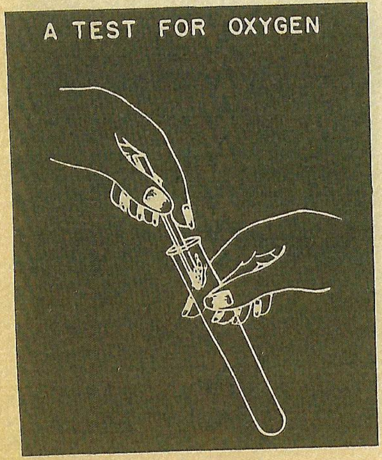
Obtain a quart milk bottle or jar and put a bunch of green leaves in it. Fill the jar full of water, placing a small piece of cardboard over the mouth and carefully turn it upside down in a pan of water so that the mouth of the jar is under the water in the pan. Carefully remove the cardboard. As long as you keep the mouth of the jar below the level of the water in the pan, no air will get into the jar or bottle. Place the pan with the bottle or jar and leaves in the sunlight and notice that a gas is given off which collects above the water in the bottle or jar. This gas is oxygen.

Try this experiment again, but keep the jar or bottle in the dark and you will find that no oxygen is given off. In this case, some carbon dioxide is given off by the leaves, but this dissolves in the water and does not collect at the top of the bottle.

EXPERIMENT 71—Making Oxygen by a Chemical Reaction.

If you can obtain some hydrogen peroxide, which is sometimes found in the medicine cabinet, you can perform this very interesting experiment. Fill a test tube ¼ full of hydrogen peroxide. In another test tube ¼ full of water, dissolve 2 measures of Sodium Carbonate (No. 4) and add 4 or 5 drops of Cobalt Chloride Solution (No. 27). Shake the tube vigorously. Pour a few drops of this pinkish liquid into the hydrogen peroxide in the other test tube. It is best to hold the test tube over a dish or over the sink while you mix the liquids because they foam up and may spill over the top. The gas being driven off is oxygen. This oxygen is obtained from the decomposition of the hydrogen peroxide. The cobalt chloride and the sodium carbonate together form basic cobalt carbonate, which helps to decompose the hydrogen peroxide without being itself consumed in the reaction. A substance which assists a reaction in this way is called a "catalyst".

$$\underset{\text{(hydrogen peroxide)}}{2\,H_2O_2} \rightarrow \underset{\text{(water)}}{2\,H_2O} + \underset{\text{(oxygen)}}{O_2\uparrow}$$



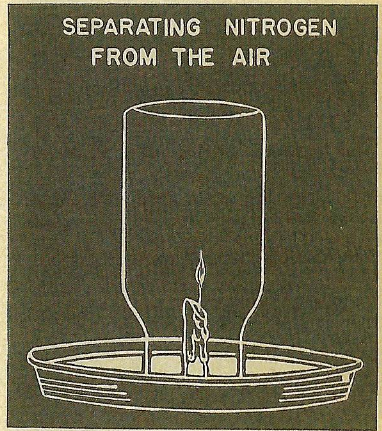
EXPERIMENT 72—A Test for Oxygen

Obtain a wooden splint about 6 inches long, ¼ inch wide and very thin. In the preceding experiment, while oxygen is being generated in the test tube, light the wooden splint at one end and after a few seconds blow it out so that a glowing spark remains on the end. Quickly thrust this glowing spark into the test tube in which oxygen is being generated and the spark will glow much brighter or perhaps burst into flame. This is the usual test for oxygen.

NITROGEN

Nitrogen is a colorless, odorless gas, which, as you have already learned, makes up nearly four-fifths of the air. This element is of great importance because it is needed for the growth of plants and it also has many uses in industry. Nitrogen is so inactive, however, that it tends to stay unchanged in the atmosphere, and even though the world is surrounded with an unlimited supply of nitrogen, it is impossible, except on a very limited scale, to make use of this element in many important ways. Early in this century, a chemist named Fritz Haber invented a process in which nitrogen of the air is made to combine with hydrogen to form ammonia. The Haber process has given the world an unlimited supply of nitrogen compounds at low cost, thus assuring ample fertilizer as well as all the nitrogen compounds we need for industrial purposes.

EXPERIMENT 73—Separating Nitrogen from Air.


SEPARATING NITROGEN FROM THE AIR

Repeat the former experiment, "Removing Oxygen from Air with Candle Flame". The gas remaining in the bottle or jar is mostly nitrogen. Place a small glass plate or piece of cardboard over the mouth of the bottle while it is still under water, and then turn the bottle right side up, keeping the cover in place so that the gas cannot escape.

EXPERIMENT 74—Wood Will not Burn in Nitrogen

Obtain a thin splinter of wood about 6 or 8 inches long and light one end of it. Remove the cover from the bottle containing nitrogen separated in the preceding experiment and thrust the burning splinter into the bottle. The flame will go out almost immediately, showing that nitrogen does not support combustion.

COMPOUNDS OF NITROGEN

The compounds of nitrogen are among the most interesting and valuable known to the chemist. With hydrogen and oxygen, nitrogen forms nitric acid. Since nitrogen is essential to the growth of plants, we find compounds of nitrogen present in almost all fertilizers. The products of animal life also contain large amounts of nitrogen. As we have learned in the Haber process, the nitrogen of the air is made to combine with hydrogen to form the compound ammonia. Ammonia is extensively used as a refrigerant, in making artificial ice, and in cold storage plants. It is used in the production of nitric acid and many other important compounds containing nitrogen.

EXPERIMENT 75—Making Ammonia.

Place 2 measures of Ammonium Chloride (No. 9) and 2 measures of Calcium Oxide (No. 20) in a test tube. Notice that neither of these chemicals has any odor before they are mixed together. Gently heat the mixture over a candle flame and after a moment, smell cautiously. Do you recognize the odor of ammonia? Moisten a piece of red Litmus Paper (No. 18) and hold this in the mouth of the test tube in which ammonia is being made. Does the red litmus paper turn blue? This simple experiment illustrates one of the ways in which ammonia is often made in chemical laboratories.

$$\underset{\text{(ammonium chloride)}}{2\,NH_4Cl} + \underset{\text{(calcium oxide)}}{CaO} + \text{heat} \rightarrow \underset{\text{(calcium chloride)}}{CaCl_2} + \underset{\text{(ammonia)}}{2\,NH_3\uparrow} + \underset{\text{(water)}}{H_2O}$$

EXPERIMENT 76—Another Method of Making Ammonia.

Place 2 measures of Ammonium Chloride (No. 9) in a test tube as before, but this time add 2 measures of Sodium Carbonate (No. 4). Heat the mixture gently, remove the test tube from the flame and after a few seconds notice the smell of ammonia at the mouth of the test tube.

EXPERIMENT 77—Making Ammonia from Organic Matter.

Obtain some small bits of horn (trimmings from the fingernails,) or a small quantity of hair. Place this in a test tube together with 2 measures of Calcium Oxide (No. 20) and add 2 or 3 drops of water. Heat this mixture gently and see if you can detect the odor of ammonia at the mouth of the test tube after removing the test tube from the flame. Moisten a small strip of red Litmus Paper (No. 18) and hold this at the mouth of the tube. If ammonia is present, the red litmus paper will be turned blue, but the color will fade away. Most organic matter contains nitrogen and when heated with calcium oxide, ammonia is formed.

EXPERIMENT 78—Making Ammonia in Your Hand.

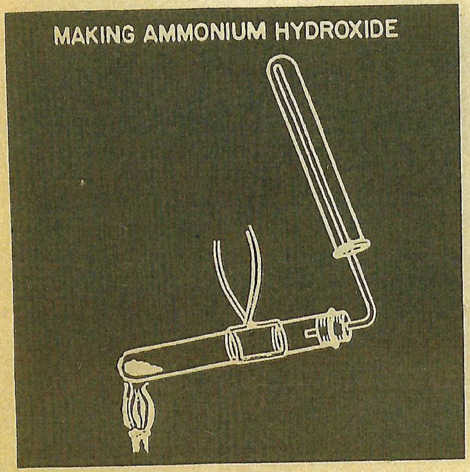
Put 1 measure of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9) in the palm of your hand and mix the substances together with your finger. Smell the mixture occasionally and notice that ammonia gas is being given off. Be sure to wash your hands immediately after performing this experiment.

EXPERIMENT 79—The Solubility of Ammonia.

Ammonia gas is very soluble in water and the liquid or household ammonia, which is commonly used, is a solution of this gas in water. Place in a test tube 2 measures of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9). Heat this mixture gently, keeping the test tube partly closed with your thumb. After you smell ammonia gas at the mouth of the tube, remove it from the flame and close the tube tightly with your thumb. Turn the tube upside down and hold the mouth under the surface of a pan of water. Now remove your thumb from the mouth of the tube and you will find the water will rise part way up inside the tube due to the ammonia dissolving in the water.

EXPERIMENT 80—Making Ammonium Hydroxide.

Put 2 measures of Calcium Oxide (No. 20) and 2 measures of Ammonium Chloride (No. 9) in a test tube. Fit the test tube with the gas delivery tube and stopper. Hold the test tube with a test tube holder and heat it over a flame with the delivery tube pointed upward. When you can smell ammonia gas coming from the end of the tube, place over it a dry test tube. Since ammonia is much lighter than air, it will rise into this second test tube and push down the air which was originally in it. Continue heating until the odor of ammonia is strong around the mouth of the inverted test tube. Then remove the test tube from the delivery tube, keeping it upside down. As quickly as possible place your thumb over the test tube mouth to prevent the ammonia from escaping. Turn it right side up, still covered by your thumb. Then remove your thumb just a little and fill the test tube ⅛ full of water. Replace your thumb and shake. You will feel a suction on your thumb as the gas dissolves in the water, forming ammonium hydroxide in accordance with the following equation:


MAKING AMMONIUM HYDROXIDE

$$\underset{\text{(ammonia)}}{NH_3} + \underset{\text{(water)}}{H_2O} \rightleftarrows \underset{\text{(ammonium hydroxide)}}{NH_4OH}$$

The reaction in this case is written with a double arrow to show that it is reversible, that is, ammonium hydroxide can be decomposed again to give off ammonia gas. Drop a small piece of red Litmus Paper (No. 18) into the ammonium hydroxide which you have just made. Does the paper turn blue?

EXPERIMENT 81—Solvent Action of Ammonium Salts.

If you have some Epsom salts (magnesium sulfate) in your medicine cabinet, dissolve 1 measure of this compound in a test tube ¼ full of water and add a few drops of household ammonia, or use the ammonium hydroxide which you prepared in the preceding experiment. Shake the test tube a little and the liquid will become cloudy with a suspension of magnesium hydroxide. Now dissolve 2 measures of Ammonium Chloride (No. 9) in a test tube ¼ full of water and add this solution to the magnesium hydroxide. You will find upon shaking the test tube that the liquid becomes clearer for the ammonium chloride dissolves the precipitated magnesium hydroxide. The solvent action of ammonium salts upon many hydroxide precipitates is important in analytical chemistry.

Ammonium salts are quite volatile and some of them pass into the form of vapor without first melting.

EXPERIMENT 82—Volatilization of Ammonium Chloride.

Put 1 measure of Ammonium Chloride (No. 9) into a clean, dry test tube. Heat gradually over a candle flame, being careful not to get the bottom of the test tube sooty. Notice that the ammonium chloride gradually disappears or turns into a vapor and part of this vapor condenses again on the cool upper part of the test tube.

EXPERIMENT 83—Cobalt Ammonia Green.

Fill a test tube ¼ full of water and add 1 measure of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9). Warm the tube over a flame for a few minutes until the odor of ammonia gas is noticeable at the mouth of the test tube. Be sure to remove the tube from the flame before bringing your face near to smell the gas. Fill another clean test tube ¼ full of water and add 3 or 4 drops of Cobalt Chloride Solution (No. 27). Now pour in just a few drops of the solution containing ammonia and you should obtain a very pretty bluish green color. This is one of the complex cobalt ammonia compounds.

EXPERIMENT 84—Copper Ammonia Blue.

Put 1 measure of Calcium Oxide (No. 20) and 1 measure of Ammonium Chloride (No. 9) in a test tube. Fill the tube ¼ full of water and warm gently over a flame until you notice a slight odor of ammonia at the mouth of the tube. In another test tube, put ½ measure of Azurite (No. 39) and fill the tube ¼ full of water. To this add a few drops of the ammonia solution from the first test tube. Close the mouth of the tube and shake so that the azurite mixes thoroughly with the liquid. Now allow this tube to stand for a few moments and as the excess azurite settles, you will notice that the clear liquid above it is a deep blue color. This color results from a soluble copper ammonia compound which is generally referred to as copper ammonia blue.

The Water On Our Earth

Two elements, which are gases at ordinary temperatures, combine to form a compound which is a liquid. These gases are hydrogen and oxygen and the compound which they form is water. Two atoms of hydrogen unite with one atom of oxygen to form this compound and the chemical formula is H_2O.

Water is the most abundant and also the most important liquid in the world. As you no doubt know, over two-thirds of the surface of our earth is covered with liquid water or with ice. The air which surrounds our earth contains a large amount of water vapor. Sometimes, when the humidity is high, and the air is warm, there may be as much as 4% by volume of water vapor in the air; and at other times, when it is quite dry, there may be almost no water vapor in the air. A man who weighs 140 pounds contains about 100 pounds of water. Our foods contain large amounts of water. Bread, for example, is almost 35% water and tomatoes will run as high as 94% water.

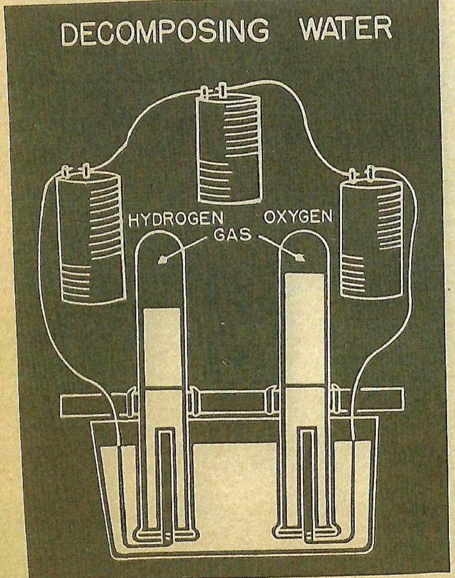
Water is necessary to all forms of plant and animal life. In the case of plants, the nutrient or plant food elements in the soil must be dissolved in water before they can be carried to the tissues of the plants where they help build additional plant cells. In our bodies, water acts as a solvent for our foods and also aids in the elimination of waste materials, helps to regulate body temperatures and acts as a lubricant.

In industry, water enters into many chemical reactions, but its most important use is as a solvent or carrier. Substances, when dissolved, are brought into much more intimate contact and react more readily than in the solid form.

EXPERIMENT 85—Producing Water from Hydrogen and Oxygen.

Fasten a candle in an upright position on a tin pan or other piece of metal, light it and hold over the candle, mouth downward, a cold, dry glass, such as a jelly glass or tumbler. Hold the mouth of the glass just above the flame. After a few seconds, notice the mist of tiny drops of water which is formed on the inside of the glass. Candle wax contains hydrogen and carbon combined in compounds called hydrocarbons. When the candle burns, the hydrogen of the hydrocarbons combines with the oxygen in the air, forming water which, due to the heat of the flame, is in the form of steam and is therefore invisible. To see this water, it must be condensed on a cold, dry object, such as the tumbler which you used.

EXPERIMENT 86—Decomposing Water into Hydrogen and Oxygen.



To perform this experiment you will need some materials that are not included in your Chemcraft outfit. Obtain 2 or 3 good dry cell batteries, 3 or 4 feet of (small) insulated copper wire, and a shallow glass or porcelain dish. Remove the carbon rods from 2 old flashlight batteries. To do this, grasp the ends with a pair of pliers, turn the carbon to break the sealing wax, and then pull the rod straight out. Fill the shallow dish about ½ full of water. Add about 5 measures of Sodium Bisulfate (No. 7) for each glass of water used. Fill 2 test tubes with this solution, close them with your thumb and then turn them upside down in the dish, being careful not to let any air get into them. As long as the mouths of the tubes are kept beneath the surface of the solution they will remain filled with the liquid.

Connect the three dry cells in "series," that is, connect the carbon binding post of one dry cell to the zinc binding post of the next dry cell. Attach long pieces of wire to the free zinc and carbon binding posts of the battery series. Fasten the free ends of these wires to the carbon rods which you have removed from the flashlight batteries, first scraping the insulation from several inches of the wire that is wrapped around the carbon rods. When this has been done, put one of the carbons under each of the test tubes, being very careful not to let any air get into either of the tubes. You will notice almost immediately that bubbles of gas form around the pieces of carbon and rise to the tops of the test tubes, displacing the water downward. In case the bubbles form very slowly or not at all, it will be necessary to connect more dry cells in the circuit in order to increase the flow of current.

Watch the gases as they collect in the tops of the test tubes. You will find that gas collects just twice as fast in one test tube as it does in the other. The water is being decomposed into hydrogen and oxygen. Water contains two atoms of hydrogen to every atom of oxygen. Hence when water is decomposed just twice as great a volume of hydrogen is given off as oxygen. The sodium bisulfate was added to make the water conduct the electric current more readily.

EXPERIMENT 87—Testing the Gases Obtained from Water.

After you have collected a considerable amount of gas in the test tubes of the preceding experiment, remove the carbons from the test tubes, being very careful to keep the mouths of the tubes under water so that no air can enter. Put your thumb over the mouth of the tube containing hydrogen and remove it from the water. Turn it right side up, bring its mouth up to a flame, and remove your thumb. As the hydrogen takes fire, you will see a pale blue flash at the mouth of the tube, accompanied by a sharp pop. This is the characteristic behavior of hydrogen when ignited in the presence of air.

In a similar way remove the tube of oxygen from the pan of water. Have ready a burning splinter several inches long, and when you are ready to remove your thumb from the mouth of the tube, blow out the flame so that only a glowing spark remains. Plunge this spark into the oxygen gas, and the splinter will burst into flame. This is the simplest test for oxygen.

THE PHYSICAL PROPERTIES OF WATER

Depending upon the temperature, water exists in three physical forms. It is a liquid at ordinary temperatures, when cooled to 32° F. it becomes a solid (ice), while when heated to 212° F. it boils and passes into the form of vapor or gas (steam).

When water boils, its vapor tension or pressure is equal to atmospheric pressure and the liquid tends to vaporize completely. It is interesting to notice that water boils more easily at high altitudes than at places near the sea level. This is because the atmospheric pressure is less at high altitudes. Since water boils at a lower temperature at high altitudes, it takes longer to cook foods than at places near sea level, because water can never be heated above its boiling temperature unless it is under pressure. Everyone is familiar with the melting of ice to form water and the boiling of water to form steam. It is interesting to remember to what a large extent our present state of civilization depends upon these physical properties of water.

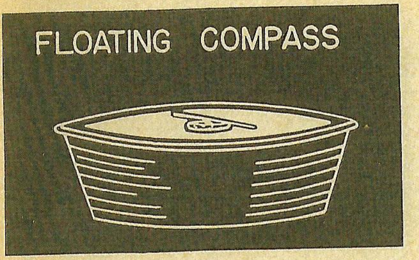
SURFACE TENSION

All liquids behave more or less as if their surface were covered with an elastic film. This arises from a force known as surface tension, which is due to the attraction between the molecules of the liquid. In the bulk of the liquid this force is exerted equally in all directions, but at the surface it makes itself felt as a contracting tendency, seeking to reduce the surface area. It is surface tension which causes a drop of liquid to take on a spherical form. It is an extremely important phenomenon in the study of emulsions and of other suspensions of minute particles. The movement of soil moisture, and of sap in plants, and the concentration of ores by the flotation process depend in large part upon surface tension.

EXPERIMENT 88—The Surface Tension of Water Supports a Safety Razor Blade.

Just as you floated a needle on water in a previous experiment, you can float a razor blade. Simply lay it gently with dry fingers on the clean surface of water in a bowl. A floating razor blade will sometimes hold up a weight of as much as a gram without sinking. You will find it interesting to touch the water at one side of the blade with a piece of soap and watch the blade be pulled away by the greater surface tension on the opposite side.

EXPERIMENT 89—A Floating Compass.



Rub a sewing needle across the pole of a magnet several times, always in the same direction. This will magnetize it. Now place it carefully on the surface of water in a bowl, and it will float. Watch it as it swings around so as to point north and south. Your compass will be a little more stable if you float it on a thin disc or cork instead of directly on the water.

EXPERIMENT 90—A Surface Tension Boat.

Make a small flat boat 2 or 3 inches long from light wood, such as cigar box lumber. Make a narrow "V" shaped slit in the stern of the boat. Insert a small piece of soap in the "V" shaped slit and carefully place your boat on the surface of clean water in a bath tub. The reduction in surface tension around the soap at the stern is so much that the tension at the forward end is able to pull the boat forward with surprising speed.

EXPERIMENT 91—Weakening the Surface Film with Soap.

Partly fill a saucer or small bowl with water, and add several measures of Sodium Carbonate (No. 4), stirring with your glass rod until the chemical is dissolved. This will soften the water so that its hardness will not destroy the soap you are going to add. If your water is very hard, it may require still more sodium carbonate.

Now float a needle or safety razor blade on the water as before, and touch a piece of soap to the surface, first on one side of the object, then on the other. Very soon the surface film will be weakened so much that the object will break through and sink.

EXPERIMENT 92—Weakening the Surface Film with Oil.

Put a drop of vegetable oil, such as cottonseed oil, on the surface of water in a saucer or bowl and see whether you can make an object float upon it. You will find that the film is now too weak to support the object, since the oil has lowered the surface tension.

This is why oil is spread on the surface of stagnant pools to prevent the breeding of mosquitoes. The larvae, wigglers, live in water, but must stick their breathing tubes through the surface to get air. Ordinarily they hang on the surface film, but when there is oil on the water, the surface tension is lowered so greatly that the surface film can no longer support them, and they sink to the bottom and die.

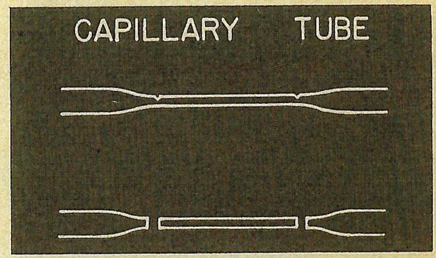
EXPERIMENT 93—The Surface Tension of Glass.

Cut a 3-inch piece from one of your glass tubes by making a scratch with a three-cornered file and then applying pressure on each side of the scratch with your thumbs and fingers. The cut end of the glass has very sharp edges. Hold it in the flame of your alcohol lamp, and when it begins to glow, you will see that the edges slowly become rounded and smooth.

Heating the glass makes it soft enough so that the surface tension can pull it into a rounded shape like a drop of liquid. Surface forces are present even in rigid solids, but they are not able to produce a change of shape until the solid is softened by heat.

Another effect of surface tension is to pull a liquid into narrow openings. If the liquid adheres to the material of which a narrow tube is made, that is, if it wets it, it will tend to creep along the surface. This increases the surface area of the liquid. Now the surface tension is always on the lookout for a chance to decrease the amount of surface, and so it tries to drag the main body of the liquid along after the creeping surface film. That is why the liquid penetrates into the tube. This is called capillary attraction.

EXPERIMENT 94— How Water Rises in a Narrow Tube.



Hold the middle of a piece of glass tube several inches long in the flame of your alcohol lamp until it softens and colors the flame yellow. Rotate the tube while you are heating it so that it will soften uniformly on all sides. Now remove it from the flame and while it is still soft pull it out several inches, constricting the central softened part to a narrow tube.

After your drawn-out tube has cooled, cut out the narrow part by scratching it near the end with a three-cornered file and breaking it. Hold the fine tube upright and dip the lower end into water in a dish. Notice that the water rises in the tube far above the level in the dish. The height of rise is greater the smaller the tube and the greater the surface tension. This is one of the ways of measuring the surface tension of liquids.

EXPERIMENT 95—Rise of Soap Solution in a Narrow Tube.

Soften half a glass of water by adding several measures of Sodium Carbonate (No. 4) or a little washing soda to it. Stir soap shavings into it until you get a good lather. Now dip into this soap solution the same end of the same tube which you used in the preceding experiment and immerse it to the same depth. Notice that the soap solution does not rise nearly so high in the tube as the water did. This is because the soap has decreased the surface tension of the water and it is therefore unable to pull itself up so far into the tube.

Many other substances besides soap have the effect of lowering the surface tension of water, and some of them are of great practical importance because the lower the surface tension of the water, the easier it is for it to wet solid surfaces. For example, in spraying leaves and fruit to kill insects and fungi, the farmer wants the liquid to stick to the leaf or fruit as a thin film, rather than to run off in drops. He accomplishes this by including in the spray mixtures a substance to lower its surface tension, thus decreasing the tendency of the liquid to gather into drops.

EXPERIMENT 96—Water Rises in a Lump of Sugar.

Put 1 or 2 teaspoonfuls of water in a saucer and add a few drops of ink. Stir this mixture until the water is uniformly colored. Now set a sugar lump on end in the colored solution, and watch how quickly the liquid rises through the sugar, being drawn up through the tiny spaces between the sugar grains just as the water was drawn up in the narrow tube used in the preceding experiment. After a time, of course, the extremely small crystals which cement together the large grains will dissolve out and the lump will crumble.

The spreading of liquids through blotting paper or a sponge is explained by surface tension in just the same way.

CONVECTION CURRENTS

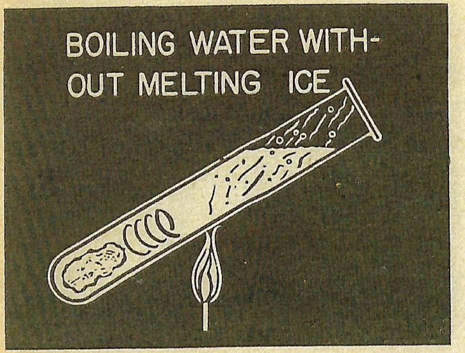
Hot water has a lower density than cold water, and so if you heat the water at the bottom of a vessel, it will rise, and the colder, denser water at the top will be forced down to take its place. You can show this very easily in a test tube.

EXPERIMENT 97—Convection Currents.

Hold the side of a test tube ¾ full of water against a flame and then add a small drop of ink. You will first notice that the color sinks to the bottom of the test tube. Then on continued heating, streamers rise up in the tube, and sink to the bottom again. A complete cycle will be described as the heated water rises, is cooled and sinks, and again rises on being reheated.

If the upper layers of water are heated instead of the lower, no convection currents can be produced, since the warm water at the top cannot rise any further.

EXPERIMENT 98—Boiling Water without Melting Ice.


BOILING WATER WITHOUT MELTING ICE

Fill a test tube nearly full of ice cold water. Obtain a small piece of ice which will just fit in the test tube, and also a weight formed from a small coil of wire which you can put on top of the ice to make it sink to the bottom of the test tube. When you have your ice in the bottom of the test tube, heat the water in the upper portion of the tube until it boils. The ice will not melt for there are no convection currents to bring warm water into contact with it.

Convection currents are possible in gases as well as in liquids, and everyone is familiar with the rising current of hot air above a radiator or stove. The winds are essentially convection currents, produced by unequal heating of different areas of the earth's surface.

Convection currents play an important part in the heating of houses, for not only is the heat distributed from the radiators to all parts of the room to a large extent by convection currents, but in the hot water type of heating system, the heat is carried from the furnace to the radiators by convection currents in the water.

TESTING WATER

All water as it occurs in nature contains more or less dissolved impurities. These are usually mineral salts, but animal and vegetable substances may also be present. Water which is to be used for domestic purposes should be examined, particularly for organic matter (that is, animal or vegetable matter). It is not that this organic matter is itself necessarily harmful but it is usually accompanied by the bacteria of putrefaction and disease.

The color, odor and taste of water are not in themselves indications of its purity, but these considered with other data help in forming an opinion of the water in question. A clear, colorless, tasteless water may be polluted, while on the other hand, a safe water may have acquired color from dissolved iron or a "peaty" taste from swamp vegetation.

EXPERIMENT 99—The Odor Test.

Fill a test tube about two-thirds full of water which you desire to test. Close the mouth of the tube with your thumb and shake thoroughly for a few moments. Now remove your thumb and notice the odor of the water. Warm the water gently over a flame and again shake and notice the odor. The odor will be strengthened by heating. Always keep your face away from a test tube while heating it and remove the tube from the flame before smelling.

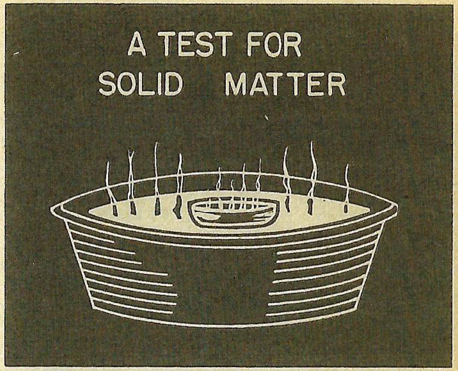
A putrid or offensive smell probably indicates sewage contamination.

EXPERIMENT 100—The Color Test.

Fill a test tube with the water to be tested, and hold the tube in front of a piece of white paper. Observe the color and clearness of the water by this method.

If it is possible to obtain a sample of distilled water, the water to be tested should be compared with the clearness of this sample.

EXPERIMENT 101—A Test for Solid Matter.


A TEST FOR SOLID MATTER

Put a small quantity of the water to be tested in a clean china or porcelain bowl and float the bowl in a pan of boiling water. The water in the bowl will gradually evaporate, and when the bowl is dry, examine carefully for any trace of solid matter, which if contained in the water will be left as a residue in the bowl.

EXPERIMENT 102—Testing for Alkalinity.

Moisten a piece of red Litmus Paper (No. 18) with a drop of the water which you are testing. If the red Litmus paper turns blue, the water is slightly alkaline.

EXPERIMENT 103—Testing for Acidity.

Moisten a piece of blue Litmus Paper (No. 18) with a drop of the water which you are testing. If the blue litmus paper turns red or pink, the water is slightly acid.

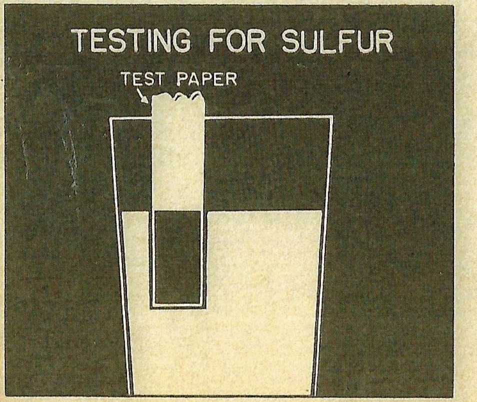
EXPERIMENT 104—Testing for Sulfates.

To a test tube of water add 1 measure of Strontium Chloride (No. 10) and shake it until dissolved. If the water, after 10 minutes, shows a white turbidity, it contains a considerable amount of sulfates, generally calcium or magnesium sulfate. Such water is usually hard.

EXPERIMENT 105—Testing for Iron.

To a test tube full of water add 1 measure of Sodium Ferrocyanide (No. 6) and shake it until dissolved. If the water shows a slight blue coloration either at once, or after standing awhile, iron is present.

EXPERIMENT 106—Testing for Sulfur.



Sulfur is present in some mineral waters in the form of hydrogen sulfide. It may be detected even in very small quantities by dropping into the water a small piece of Sulfide Test Paper (No. 31). If hydrogen sulfide or other sulfides are present, the paper will turn brown or black.

EXPERIMENT 107—Testing for Carbon Dioxide or Carbonates.

Make up a solution of lime water by adding ½ measure of Calcium Oxide (No. 20) to ½ test tube of water. Shake the tube well and let it settle for a few minutes. Add a few drops of the clear portion of this solution to a tube of the water which is to be tested. If a white turbidity or precipitate is formed, the water contains carbon dioxide, either as a dissolved gas or in the form of a carbonate or bicarbonate.

EXPERIMENT 108—Testing Water for Calcium and Magnesium.

To a test tube half full of the water to be tested add 1 measure of Sodium Carbonate (No. 4) and shake the tube to dissolve the solid. If, after standing a little while, the water shows a white turbidity or a precipitate, calcium or magnesium salts are present.

Now add 2 measures of Ammonium Chloride (No. 9) and shake the tube again. If the white precipitate which was formed by the sodium carbonate is dissolved by the ammonium chloride, the original water contained magnesium. If the precipitate does not dissolve, calcium was present, probably in the form of calcium sulfate or calcium bicarbonate. Ordinarily, hard water contains considerable quantities of calcium salts accompanied by a lesser amount of magnesium salts.

HARD WATER

The hardness of water is caused by dissolved calcium and magnesium salts, which react with soap to convert it into insoluble and useless compounds. The waste of soap caused by hard water amounts, in some communities, to many thousands of dollars a year. In steam boilers, and in heating systems, the salts in hard water tend to precipitate out as a hard scale, coating the inside surfaces of boilers and pipes. This hinders the transmission of heat and the flow of the liquid.

There are two general types of hardness which may be present in water. The first is bicarbonate hardness, caused by calcium or magnesium bicarbonates. These two compounds are converted by boiling into calcium or magnesium carbonate, substances which, being insoluble, precipitate out of the water. Therefore, bicarbonate hardness can be removed by boiling the water. For that reason it was formerly called temporary hardness.

Non-carbonate hardness is the second type, and is caused by calcium or magnesium sulfates. Since these compounds are not precipitated by boiling, water containing them was formerly called permanently hard. Both types of hard water, however, are very readily softened by certain chemicals. In the household, washing soda (sodium carbonate), ammonia water (ammonium hydroxide), and trisodium phosphate are often used. All of these precipitate the calcium and magnesium in the water by converting them into insoluble compounds.

A better softening method makes use of a mineral known as zeolite. This substance is a sodium aluminum silicate, and when hard water comes in contact with it, calcium and magnesium ions dissolved in the water are exchanged for the sodium of the mineral and are held fast on the surface of many tiny pores of the solid. Thus, when hard water containing calcium sulfate is allowed to trickle through a bed of zeolite, the calcium sulfate is converted into sodium sulfate, which does not react with soap. Hence, the water is softened.

Of course there is a limit to the amount of calcium and magnesium which the zeolite can remove from water by exchange with sodium, but when its exchange capacity is exhausted, it is a simple matter to restore its activity. This is done simply by washing it with a concentrated salt solution, which puts sodium ions back into the zeolite and removes the calcium and magnesium which the mineral has taken up. After the excess salt has been washed out, the zeolite is back in its original state and can be used to soften another quantity of water. This cycle of use and regeneration can be repeated indefinitely, and the zeolite never needs to be replaced.

A more recent development in water treatment uses various specially designed synthetic resins. These are chemically very different from the zeolite, but they have the same property of exchanging one kind of ion for another. For ordinary water softening they are made to remove calcium and magnesium ions from solution, replacing them with sodium ions. Like the zeolite water softeners, these resins can be used over and over again by simply regenerating them with a concentrated salt solution. This method of water softening is now used in many homes.

Other resins of the same general type will do even better than this and remove from water nearly all ions, both positive and negative. The result is an extremely pure water, such as formerly could be obtained only by the expensive process of distillation. Such ion exchange resins are being used more and more to supply pure water for industrial processes and chemical laboratories.

EXPERIMENT 109—The Effect of Hard Water on Soap.

Make some artificial hard water by adding ½ measure of Epsom salts (magnesium sulfate) to a test tube half full of water. Shake the tube until the solid is dissolved. Now make some soap solution by shaking 2 measures of fine soap shavings in a test tube half full of warm water. Add the soap solution drop by drop to the hard water and notice the curdy precipitate which forms. This is a magnesium soap, which is insoluble in water. It is of no use for washing.

EXPERIMENT 110—Testing Water for Hardness.

Add 2 measures of fine soap shavings to a test tube half full of water and after shaking it well, warm gently to dissolve the soap completely. Fill another test tube half full of the water to be tested and add to it 2 or 3 drops of the soap solution. Close the test tube with your thumb and shake it vigorously. If a permanent foam or suds forms, the water is soft. If not, add a few drops more of soap solution and shake the tube again. Continue this until a foam is produced. The amount of soap solution required to produce a foam is an indication of the hardness of the water. Try this with rain water collected in a clean dish.

EXPERIMENT 111—Softening Hard Water.

Prepare some artifically hard water by shaking ½ measure of Calcium Oxide (No. 20) in a test tube ⅔ full of water. Let the undissolved solid settle to the bottom of the tube and pour off a little of the clear liquid into another test tube. To this portion add several drops of soap solution prepared by shaking a few soap shavings with water. Shake the tube vigorously after you add the soap solution and you will find that no suds are formed. If you continue adding soap solution until suds form when the tube is shaken, you will realize that the water was very hard since a large amount of soap solution was required. Furthermore, you will notice a curdy precipitate of calcium soap.

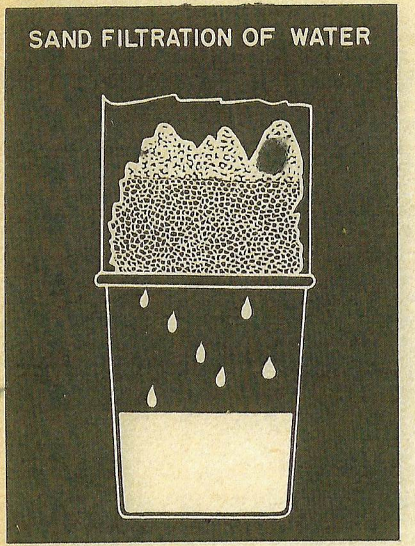
Now pour another portion of your artificially hard water into another test tube and add 1 measure of Sodium Carbonate (No. 4). Shake the tube and notice the turbidity caused by the calcium carbonate which is precipitated. The calcium hardness has now been removed from the water and when you add a few drops of soap solution you will find upon shaking the tube that a good lather or suds forms at once.

FILTERING AND CLARIFYING WATER

Many natural waters in addition to dissolved impurities contain some matter in suspension. Most cities which must use muddy water make use of a filtration process to clarify their water.

While the compounds causing hardness are dissolved and must be removed by chemical means, suspended particles which make water appear cloudy or muddy can be removed by filtration. This is done on a large scale by allowing the water to percolate through beds of sand.

EXPERIMENT 112—Filtering Water Through Sand.



Obtain a small, clean tin can and with a nail, punch about ten small holes in its bottom. Now fill it nearly full of fine clean sand. It may be well to wash the sand first by stirring it up in a pail of water and allowing it to settle. Set the can of sand on top of a clean tumbler.

Get some muddy water, or if need be, make it by stirring up a little clay or earth with water. Pour this into your sand filter, and soon clear water will begin to drip from the bottom.

Often a chemical method of clarification is applied before the water is filtered. Aluminum sulfate is added in small quantities and if the water is not sufficiently alkaline to precipitate aluminum hydroxide directly, a little lime or sodium carbonate is needed also. The hydroxide formed is a flaky, gelatinous precipitate and as it settles out, it entangles and drags down with it suspended particles and bacteria, leaving the water much clearer.

EXPERIMENT 113—Clarification of Water by a Chemical Precipitate.

To a test tube nearly full of muddy water (made so if necessary by the addition of a small amount of clay) add ½ measure of Aluminum Sulfate (No. 12) and shake the tube until the solid is dissolved. Next add ½ measure of Sodium Carbonate (No. 4) and shake it again. Let stand quietly for 15 minutes. For comparison let some of the untreated muddy water stand in another test tube along side.

SOLUTIONS

When we put sugar in tea or coffee, when we wash our hands with soap and water, when we add salt to water for cooking vegetables, we are making solutions. Practically every liquid we use is a solution. Even tap water contains not only mineral matter but also gases dissolved in it.

A substance is dissolved or "in solution" when it is so finely divided, and its particles are so completely scattered throughout the liquid that they can neither be seen nor separated from the liquid by filtering. A dissolved substance may in most cases be recovered in its original state by evaporating the liquid.

EXPERIMENT 114—Making a Solution and Recovering the Dissolved Substance.

Dissolve a spoonful of common table salt (not included in outfit) in half a glass of water. Now pour this solution into a small pan or similar cooking utensil and heat on the stove until the water is driven off. Taste what you find remaining in the pan and satisfy yourself it is the same substance that you started with.

Unless the pan is heated very gently when the salt is nearly dry you will find there will be some popping due to the expansion of steam in the crystals of salt. This phenomenon is known as decrepitation. It does no harm except to scatter some of the salt.

EXPERIMENT 115—An Insoluble Substance.

Fill a test tube half full of water and add 1 measure of Sulfur (No. 1). Close the mouth of the tube with your thumb and shake it thoroughly. The sulfur will not dissolve but remains in a solid state scattered throughout the water.

To prove that none of the sulfur has dissolved in the water, let the contents of the test tube settle and put some of the clear liquid in a clean, tin cup and heat very gently until it has entirely evaporated or boiled away. There will be no solid remaining as when you evaporated the salt solution.

Most chemicals react much more readily if they are dissolved in water, as this permits the particles or molecules to come into closer contact.

EXPERIMENT 116—Separating a Soluble from an Insoluble Substance.

Place in a test tube 3 measures of Ammonium Chloride (No. 9) and 2 measures of Sulfur (No. 1) and shake the tube until they are well mixed. In order to separate them again, fill the test tube two-thirds full of water. Since the ammonium chloride is soluble in water, a little shaking will dissolve it, but the sulfur is insoluble and remains as a solid suspended in the solution of ammonium chloride.

Now filter this mixture through a piece of soft paper as described in "The Use of Chemical Apparatus." If you have worked carefully, the filtrate will be clear and free from suspended sulfur. Pour it into a saucer and set it in a warm place. When the water has entirely evaporated you will find remaining a white deposit of ammonium chloride. Remove the filter paper and spread it out to dry. On it you will find the sulfur you originally mixed with the ammonium chloride.

EXPERIMENT 117—Effect of Solution on Chemical Reaction.

Mix together on a sheet of paper 1 measure of Sodium Carbonate (No. 4) and 2 measures of Sodium Bisulfate (No. 7). The dry powder obtained shows no signs of any chemical reaction. Now put this mixture in a test tube and add a drop or two of water. Immediately a reaction occurs liberating a large quantity of gas. This gas is carbon dioxide.

EXPERIMENT 118—Another Example Showing Color.

Mix together on a sheet of paper 1 measure of Sodium Ferrocyanide (No. 6) and 1 measure of Ferric Ammonium Sulfate (No. 21). Here again the dry powder obtained shows no sign of any chemical reaction. Put this mixture in a test tube and fill the test tube ¼ full of water. Immediately an intense blue color is formed showing that the two substances have reacted forming still another substance.

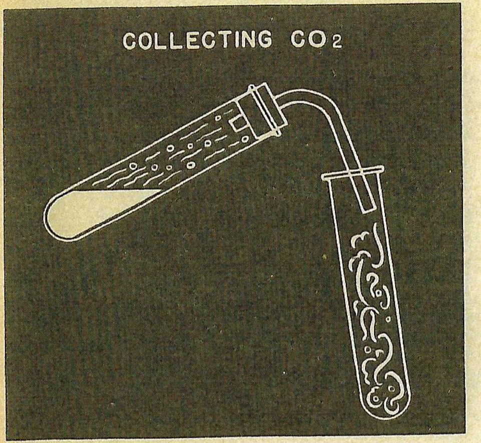
The particles of a dissolved substance have the property of spreading themselves uniformly throughout the entire volume of a liquid. This scattering is known as diffusion.

EXPERIMENT 119—Diffusion.

To a glass ¾ full of water add 1 measure of Ferric Ammonium Sulfate (No. 21) and stir until it is completely dissolved. Now get a piece of tissue paper about 2 inches square and in its center, place 1 measure of Sodium Ferrocyanide (No. 6). Fold up the solid tightly in the paper and twist the corners together. Put the paper wad in the glass of water and set it in a quiet place. When the liquid soaks through the paper, blue streamers will descend from it, and eventually the forces of diffusion will make the liquid a uniform blue color.

Gases as well as solids tend to dissolve in liquids. It is oxygen dissolved in water which makes it possible for fish to live. Pollution by sewage and industrial wastes tends to remove the dissolved oxygen from rivers, but as more oxygen dissolves from the air, the organic waste materials are slowly oxidized, and the water is purified. A stream with many rapids and waterfalls will purify itself more quickly than one which runs quietly since oxygen will have a better chance to dissolve in it.

EXPERIMENT 120—Dissolving a Gas in Water.



Put 2 measures of Sodium Carbonate (No. 4) and 2 measures of Sodium Bisulfate (No. 7) in a test tube and fill one-quarter full of water. Quickly fit the test tube with your gas delivery tube and stopper, and insert the other end of the delivery tube in the mouth of an empty test tube so that the gas which is coming from the delivery tube will flow into the test tube just as if water were coming through the delivery tube instead of carbon dioxide gas, which you cannot see.

After a minute or so the test tube should be full of gas. Keep your finger over the mouth as much as possible and carefully fill the tube one-quarter full of water. Now press your finger or thumb tightly over the mouth and shake the tube. Do you feel a suction on your thumb showing that the gas in the tube has dissolved in the water forming a partial vacuum?

Gases always become less and less soluble the higher the temperature. Therefore, if a liquid contains the greatest possible amount of gas dissolved in it at low temperature, some of this gas must come out of the solution when the liquid is heated.

EXPERIMENT 121—The Effect of Temperature on Solubility of a Gas.

Fill a test tube half full of water drawn from the cold water faucet, and heat it slowly and gently over a flame. Notice that as the water becomes warm, bubbles of gas appear on the sides of the tube. This is air which was dissolved in the water at the lower temperature. By the time the water is nearly up to its boiling point, the sides of the tube will be thickly covered with bubbles. Boil the water for a few seconds, and then cool the tube in cold water. Again heat the water slowly and gently, and this time you will see no bubbles, or at most very few, since the previous heating removed all the dissolved air.

Now cool the test tube once more, and shake it vigorously, thereby giving some more air a chance to dissolve in it. Upon heating it again, you will find that bubbles form just as they did on the first heating.

When chemists want to remove a gas from a liquid, they usually boil the liquid.

EXPERIMENT 122—Removing a Gas from Solution by Boiling.

Place 1 drop of household ammonia in a test tube ¼ full of water. Now add 1 drop of Phenolphthalein Solution (No. 22). This will color the liquid in the test tube red, as ammonium hydroxide is a base. Heat this red liquid for a few minutes and notice that as it begins to boil the red color becomes fainter. By boiling for several minutes, the liquid may be turned colorless, but be careful not to let the hot liquid spurt out of the test tube.

Ammonium hydroxide is formed by dissolving ammonia gas in water. When the solution is boiled, the ammonia is driven out.

EXPERIMENT 123—Solubility of Liquids in Each Other.

Try to mix a few drops of light oil with a little water in a test tube. You can shake it as much as you wish, but the oil will always separate from the water. In other words, it is insoluble in water.

Now add a few drops of glycerine to a test tube ¼ full of water. As soon as you begin to shake the tube, the glycerine immediately dissolves in the water, showing that glycerine is soluble in water.

Many salts when dissolved absorb heat. Thus it is possible to lower the temperature of water by dissolving certain salts in it.

EXPERIMENT 124—Lowering Temperature by Solution.

Fill a test tube ¼ full of water and feel the temperature by holding the test tube against your cheek for a few seconds. Now add 4 measures of Sodium Thiosulfate (No. 16) and shake the tube for a moment. Bring the bottom of the tube in contact with your cheek again and notice the decided difference in the temperature.

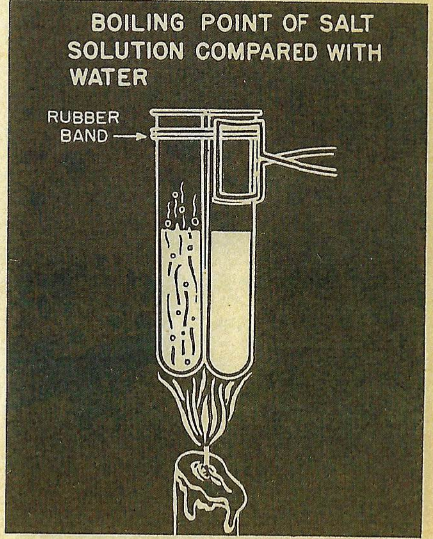
EXPERIMENT 125—Raising Temperature by Solution.

Place 5 measures of Sodium Carbonate (No. 4) in a dry test tube. Warm the tube over your alcohol lamp flame and you will notice that water of hydration is given off in the form of steam. Continue warming the tube until no more water is given off. Now set this tube aside to cool. After about 15 minutes when the tube has become cool, fill another test tube ¼ full of water and carefully notice its temperature by holding the bottom of the tube against your wrist. Now pour the water in this tube into the tube containing the dehydrated sodium carbonate. Close the mouth of the tube with your thumb and shake it. After the sodium carbonate has dissolved, again check the temperature of the bottom of the tube and notice if the water has become warmer by dissolving the dehydrated sodium carbonate.

Dissolving a salt in water raises the boiling point of the water. For example, a solution containing 25 parts of sodium chloride in 100 parts of water boils at 223° Fahrenheit (106° Centigrade). Pure water, of course, boils at 212° F. (100° C.).

EXPERIMENT 126—Boiling Point of Salt Solution Compared with Water.

Dissolve 10 measures of common salt in a test tube half full of water, and fill another test tube half full of plain water. Now heat these two test tubes over a flame. Hold them so that each will get just the same amount of heat. You will find that the water will boil before the salt solution does. Hold the tubes carefully so the boiling water will not spurt out on you.


BOILING POINT OF SALT SOLUTION COMPARED WITH WATER
RUBBER BAND

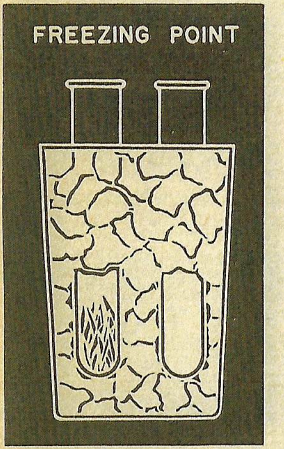
Dissolved solids not only raise the boiling point of water, but lower its freezing point as well. This explains how salt melts ice on slippery sidewalks, and why a mixture of ice and salt is used to freeze ice cream. Saying that salt lowers the freezing point of water is the same as saying that it lowers the melting point of ice. Suppose that we add salt to a mixture of ice and water which is just at the freezing temperature. The salt coming in contact with the ice lowers its melting point below 32 degrees Fahrenheit, but since the mixture of ice and water is at just 32 degrees, the ice is warmer than its melting point and so it begins to melt. In doing so it takes up heat from the water and anything else in contact with it just as a cake of melting ice takes up heat from your hand. The ice in the mixture will continue to melt and take up heat until it has cooled the whole mixture down to the freezing point of the salt solution. Of course if water or an ice cream mix is placed in this freezing mixture it will be frozen.

EXPERIMENT 127—The Action of Freezing Mixtures.

Crack enough ice to fill a glass or tumbler ¾ full. Now put about ¼ glass of salt in with the ice and mix the two together. Fill a test tube half full of clear water and set the tube in the mixture of salt and ice for a few minutes so that it is surrounded by the freezing mixture. Be careful that you do not allow any of the salt to get into the test tube. In a short time the water in the test tube will be frozen.

Now you can prove directly that dissolved salt lowers the freezing point of water.

EXPERIMENT 128—Freezing Point of Salt Solution Compared with Water.



Prepare a freezing mixture by mixing ice and salt just as in the preceding experiment. Fill a test tube half full of water and add 8 or 10 measures of common table salt. Shake the tube until the salt is dissolved. Fill another test tube half full of pure water. Now put both of these test tubes into the glass containing the mixture of ice and salt, and you will soon find that the pure water will freeze, while the salt solution will not.

Alcohol, as well as salt, lowers the freezing point of water, and for that reason it is used in automobile radiators in the winter. People do not put salt in their radiators to keep them from freezing because the salt increases the corrosion of metal.

Although a liquid ought to freeze when cooled down to its freezing point, it sometimes happens that it remains liquid even when the temperature is far below the freezing point. A substance which, although it is below its freezing point, is still liquid is said to be undercooled. An undercooled liquid is in an unstable condition, and sometimes the slightest jar will cause it to solidify.

EXPERIMENT 129—Undercooled Water.

Obtain half a glass of cracked ice and mix with it several spoonfuls of ordinary salt. Fill a test tube ¼ full of pure water and place it in the freezing mixture. Be careful not to allow any of the ice or salt to get into the test tube. If the water in the test tube is kept perfectly quiet, it may be cooled as much as 10° Centigrade below zero (that is, to 14° F.) without freezing. Then if it is shaken, or if a tiny crystal of ice is added, it will immediately begin to freeze and the temperature will rise to the freezing point.

We have learned that gases are always less soluble in hot water than in cold. A few solid substances obey the same rule, but most solids behave in just the opposite way, that is, they are more soluble in hot water than in cold.

EXPERIMENT 130—The Usual Effect of Temperature on Solubility.

Put 4 measures of Ammonium Chloride (No. 9) in a clean, dry test tube. Add exactly 4 drops of water. Shake the tube for several minutes and notice that the ammonium chloride will not all dissolve in this amount of water. Now warm this solution, shaking the tube as you warm it. You will see more and more of the ammonium chloride going into solution as the liquid becomes hot until finally, it should be practically all dissolved. Set the tube to one side and allow the solution to cool slowly. After awhile, crystals of ammonium chloride will form and will gradually grow larger.

EXPERIMENT 131—The Opposite Effect of Temperature on Solubility.

Fill a test tube half full of cold water and add ½ measure of Calcium Oxide (No. 20). Calcium oxide combines with water to form calcium hydroxide. Shake the tube thoroughly in order to dissolve all the calcium hydroxide possible. Now allow the tube to remain quiet for a minute, and then pour part of the clear solution into another clean test tube being very careful to leave the undissolved calcium hydroxide in the first test tube.

Heat the solution in the second tube over a flame for a few moments, and you will find that the clear solution gradually becomes slightly turbid or cloudy, since calcium hydroxide is less soluble in hot water than in cold water and part of it comes out of solution when the temperature is raised.

Sometimes it is possible to prepare a solution which contains more of a dissolved substance than is normal for the temperature. Such a solution is said to be supersaturated, and like an undercooled liquid, it is in an unstable state. Crystallization of the excess of the dissolved substance may be started by a jar or by dropping in the slightest trace of solid.

EXPERIMENT 132—A Supersaturated Solution.

Fill a test tube ⅓ full of Sodium Thiosulfate (No. 16). Add exactly 2 drops of water and warm gently. The sodium thiosulfate will melt, or dissolve in the water. Continue heating until the solid is entirely dissolved and the solution is almost boiling. Now close the mouth of the test tube with a plug of cotton and allow it to stand without being shaken until it is cold.

You now have a supersaturated solution, that is, one containing much more dissolved substance than it has any right to have at room temperature. Drop in a tiny crystal of sodium thiosulfate and at once crystals will begin to form in the supersaturated solution.

CRYSTALLIZATION

The process of crystallization is very important to the chemist, since by means of it he can purify many substances. When crystals form from a solution, they tend to reject the impurities, leaving them behind in the solution. Hence to purify a compound, we dissolve it in liquid and then allow it to crystallize out again. The crystals obtained by this recrystallization are purer than the original ones, the impurities having remained in the solution.

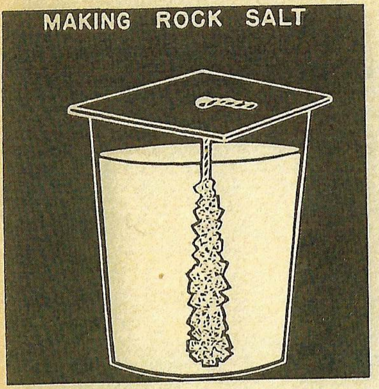
When a substance is more soluble in hot water than in cold, it is usually easy to get good crystals. All we need do is dissolve as much of the substance as possible in hot water and then let the solution cool. When cold, the solid is less soluble, so some of it must separate out as crystals. It is an invariable rule that the more slowly crystals are formed the larger they will be.

EXPERIMENT 133—The Formation of Crystals.

Put 3 measures of Sodium Ferrocyanide (No. 6) in a test tube and add 5 drops of water. Heat for a minute to dissolve the solid, then set it aside and let the solution cool slowly. After 15 or 20 minutes, a heavy deposit of yellowish crystals will be found in the test tube.

Repeat the experiment, but cool the solution rapidly as soon as the sodium ferrocyanide has all dissolved by setting the test tube in a glass of cold water. The crystals obtained this time will be much smaller than when the solution is cooled slowly.

EXPERIMENT 134—Making Rock salt.



Fill a glass tumber ¼ full of common salt and then add warm water until the glass is ¾ full. Stir the solution until the salt is all dissolved, or if necessary, add a little more water.

Cut a piece of cardboard to fit as a cover for the glass and hang a piece of string through a hole in the center of the cardboard so that it reaches nearly to the bottom of the glass. It may be necessary to tie a small weight on the string to make it sink in the solution. Now cover the glass and set it aside in a quiet place with the string hanging in the solution. Observe the glass from day to day and after a time you will notice crystals of salt forming on the string. After several days these crystals will grow quite large and you can lift them out of the glass by means of the string.

EXPERIMENT 135—Rock Candy.

Fill a glass half full of very hot water and add sugar until no more will dissolve after stirring thoroughly. Pour this clear syrup into a clean glass and fit a cover and string as in the rock salt experiment. After the sugar solution has remained standing for a few days, large crystals of rock candy will form on the string.

EXPERIMENT 136—Crystallizing A Mixture of Salts.

Place in a test tube 1 measure of Sodium Bisulfate (No. 7), 1 measure of Ferric Ammonium Sulfate (No. 21), 1 measure of magnesium sulfate (Epsom salts) and 6 or 7 drops of Cobalt Chloride Solution (No. 27). Carefully add water until the tube is about ¼ full and then heat gently until the solids are dissolved. Now pour the clear, hot solution into a clean glass, being careful to leave any undissolved solids in your test tube. Set the glass in a quiet place and cover with a small piece of cardboard. Notice after a day or two the crystals which have formed in the solution. Chemists often separate different compounds from each other by crystallization since some substances crystallize more easily than others.

EXPERIMENT 137—Frosting Glass with Crystals.

Put about ½ inch of water in a test tube and add 5 measures of Ammonium Chloride (No. 9) and only 1 small drop of liquid glue or mucilage. Now heat the solution until the solids are all dissolved and apply the hot liquid, using a small brush, to the glass which you wish to frost. The liquid will begin to evaporate almost immediately and crystals will be formed which give the glass the appearance of a frosted pane.

On a larger scale, frosting glass can be done by dissolving magnesium sulfate (Epsom salts) in hot water until no more will go into the solution. Add a few drops of liquid glue or mucilage and apply this solution to the glass to be frosted.

Water is contained in many solid dry substances in a state of chemical combination. For example, some crystals contain what is called "water of hydration" or "water of crystallization".

EXPERIMENT 138—Water of Hydration.

Drop 1 measure of Sodium Ferrocyanide (No. 6) into a clean dry test tube and heat the bottom of the tube with a candle flame. Do not hold the tube in the flame, as this will cause soot to deposit. Rather hold it about ½ inch above the top of the flame. Keep your face away from the test tube while you are heating it. Notice that the powder becomes wet and water is driven off in the form of steam.

EXPERIMENT 139—Water of Hydration in Washing Soda.

Washing soda is a good example of a substance containing water of hydration. If you are familiar with this material, you know that it sometimes comes in clear, nearly transparent, crystals while at other times it is covered with white powder, and is very dusty. This is because washing soda loses part of its water of hydration when it is exposed to the air, and the white dusty powder has probably been stored so long that much of the water has been lost. It is just as well to remember, however, that a pound of washing soda that has lost part of its water contains much more soda than a pound of the big, good-looking crystals which still have the water in them. Heat a few measures of washing soda in a test tube and notice the large amount of water which is given off.

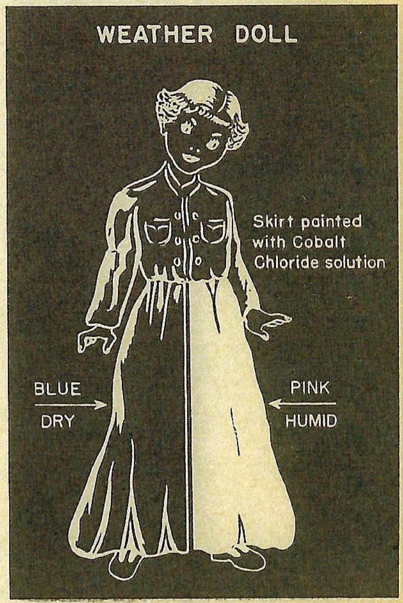
EXPERIMENT 140—Water of Hydration in Cobalt Chloride.

Put 5 or 6 drops of Cobalt Chloride Solution (No. 27) in a clean test tube and warm this solution gently. Suddenly you will notice that the pink solution has turned blue. Stop heating and allow the test tube to cool and the solution will again be pink. Heat the tube again until the moisture is all driven off from the cobalt chloride and you will find a blue solid remaining. Now add just 1 drop of water and the blue solid again becomes pink. These changes are due to water of hydration. When heated, the water of hydration is driven off from the cobalt chloride and it becomes blue in color. When allowed to cool, the cobalt chloride again takes up moisture as water of hydration and becomes pink.

EXPERIMENT 141—A Cobalt Chloride Barometer.

Cobalt chloride is often used as a form of chemical barometer. A sheet of white paper is moistened with a concentrated solution and allowed to dry. It is then hung where the outside air will circulate freely about it, but where the rain will not reach it. When it looks bluish it is an indication of dryness and usually an indication of continued clear weather. When it is pink it indicates wet or moist weather.

Such a barometer can easily be made by painting Cobalt Chloride Solution (No. 27) on a sheet of paper with a soft brush. These barometers or weather indicators are often seen in the form of dolls or figures whose dresses are painted with cobalt chloride solution, or they can be made up in the form of pictures which change color according to the weather.


WEATHER DOLL
Skirt painted with Cobalt Chloride solution
BLUE
DRY
PINK
HUMID

EXPERIMENT 142—Testing for Traces of Moisture.

A convenient method of testing for small traces of moisture is as follows: Prepare a solution of 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. Saturate some small strips of soft, white paper with this solution and dry the strips. When you are ready to use them, place half a measure of Sodium Ferrocyanide (No. 6) on one of the dry strips and spread it over the surface by means of a soft brush or wad of cotton.

The slightest trace of moisture produces dark blue spots on these strips.

HYDROGEN

Hydrogen is a colorless, odorless gas. As we have learned, it is contained in water in combination with oxygen. Practically all organic substances, such as wood, petroleum, sugar, fats and proteins contain hydrogen combined with carbon, and often with other elements. Hydrogen is an essential constituent of all the acids, and is often referred to as "the acid former".

Because it is the lightest gas known, hydrogen was sometimes used in balloons and zeppelins, but since its inflammability has led to numerous serious accidents, an example of which is the Von Hindenburg disaster, the more expensive helium has replaced it for this purpose.

Hydrogen is a rather active element, combining, for example, directly with oxygen, that is, burning, to form water. As we have already learned, ammonia is formed when hydrogen reacts with nitrogen under pressure. In the so-called hydrogenation of oils, hydrogen under pressure is caused to combine with hot vegetable oils to form solid cooking fats. A nickel catalyst assists the reaction.

One of the most common ways of preparing hydrogen in the laboratory is by the reaction between a metal and an acid. Sodium bisulfate, $NaHSO_4$, can be considered as partly sodium sulfate and partly sulfuric acid. When dissolved in water, this chemical forms the equivalent of a dilute solution of sulfuric acid. In the following experiments we will, therefore, use sodium bisulfate dissolved in water and will consider the reactions between the sulfuric acid contained in this solution and the metal.

EXPERIMENT 143—Making Hydrogen from Sulfuric Acid with Iron.

Prepare a solution of 4 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water. Warm the solution slightly, if necessary, to dissolve the solid. Now add 1 measure of powdered iron metal or small pieces of iron obtained by cutting fine steel wool with scissors. There will be a reaction and hydrogen gas is given off. In this experiment, the hydrogen in the sulfuric acid was replaced by the iron, forming ferrous (iron) sulfate and hydrogen.

$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(iron)}}{Fe} \rightarrow \underset{\text{(iron sulfate)}}{FeSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 144—Making Hydrogen with Aluminum.

Make a solution of sulfuric acid by dissolving Sodium Bisulfate (No. 7) in water as in the preceding experiment and add a small piece of aluminum metal. You can probably obtain aluminum from an old pan or kettle which is no longer good for cooking purposes. Notice that the acid reacts with the aluminum and that bubbles of hydrogen gas form on the metal.

In this experiment the hydrogen of the sulfuric acid was replaced by the aluminum, forming aluminum sulfate and hydrogen.

$$\underset{\text{(sulfuric acid)}}{3\ H_2SO_4} + \underset{\text{(aluminum)}}{2\ Al} \rightarrow \underset{\text{(aluminum sulfate)}}{Al_2(SO_4)_3} + \underset{\text{(hydrogen)}}{3\ H_2\uparrow}$$

EXPERIMENT 145—Making Hydrogen with Zinc.

Obtain a piece of zinc from an old flashlight battery and after cleaning it so that it is bright and shiny, cut off a few small pieces. Now prepare a solution of sulfuric acid as in the preceding experiment and to the sulfuric acid solution add a few of the small pieces of zinc. You will again find bubbles of hydrogen gas given off as the hydrogen of the sulfuric acid is replaced by the zinc, forming zinc sulfate and hydrogen. In this experiment the action should be more vigorous than in the case of either the iron or the aluminum.

$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(zinc)}}{Zn} \rightarrow \underset{\text{(zinc sulfate)}}{ZnSO_4} + \underset{\text{(hydrogen)}}{H_2\uparrow}$$

EXPERIMENT 146—Filling Bubble Balloons with Hydrogen.

Make some soap suds by stirring a small piece of soap in a saucer of soft water. Now prepare hydrogen as you did in the preceding experiment. Be sure to use plenty of zinc and cut it into very thin slivers or pieces. Fit your gas delivery tube to the mouth of the test tube as soon as the gas begins to come off from the zinc.

Insert the end of the gas delivery tube under the surface of the soap suds in the saucer and as the hydrogen gas emerges from the tube, bubbles will be formed that will rise through the air. Each of these hydrogen-filled bubbles is in effect a tiny ballon buoyed up by the hydrogen which it contains.

EXPERIMENT 147—Displacement of Hydrogen from an Alkali by Aluminum.

Place 3 measures of Sodium Carbonate (No. 4) and 1 measure of Calcium Oxide (No. 20) in a test tube half full of water. Shake the test tube thoroughly to dissolve all of the solids that you can. Now allow the tube to stand quietly for a few minutes and then pour off the clear liquid into another test tube. Add some small pieces of aluminum to the solution in this test tube and warm slightly over a flame. You will soon notice bubbles of hydrogen gas coming off rapidly. While the hydrogen gas is being given off, bring a flame near the mouth of the tube and you will probably notice a slight puff as the hydrogen gas burns. Be sure to keep your face at a distance when you try this.

You formed a dilute solution of sodium hydroxide when you mixed the sodium carbonate and the calcium oxide. The aluminum metal displaced part of the hydrogen from the sodium hydroxide, forming aluminum hydroxide, which is called sodium aluminate.

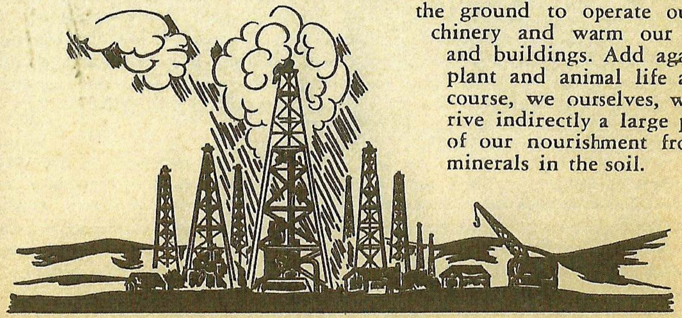
$$\underset{\text{(aluminum)}}{2\ Al} + \underset{\text{(sodium hydroxide)}}{6\ NaOH} \rightarrow \underset{\text{(hydrogen)}}{3\ H_2\uparrow} + \underset{\text{(sodium aluminate)}}{2\ Na_3AlO_3}$$

The Chemical Wealth Within Our Earth

We have learned about the atmosphere which surrounds our earth and about the water which is on our earth, so we will now take up the earth itself. What is in our earth? The surface may consist of rocks or a very thin layer of clay or sand or gravel, usually covered with a few inches of loam, or top soil. Underneath we have a variety of materials, mostly classified as minerals and rocks.

A mineral is a natural inorganic substance, usually having a definite chemical composition and often a characteristic crystalline form. Rocks are large masses of minerals or mixtures of minerals. They do not have a definite chemical composition.

Minerals are our world's greatest source of wealth. Consider the value of the iron, copper, lead, zinc, aluminum; and the precious metals, gold, silver and platinum; the diamonds and other gems. Add the value of the coal and oil, which we take from the ground to operate our machinery and warm our homes and buildings. Add again the plant and animal life and, of course, we ourselves, who derive indirectly a large portion of our nourishment from the minerals in the soil.



From a chemical standpoint, minerals may be either impure metallic or non-metallic elements, or compounds of elements. In either case, we must purify and convert them into the desired form. In the following pages we will take up many of the more common elements and learn from what minerals they are obtained and by what processes they are converted into usable form.

THE PERIODIC SYSTEM

Chemists have classified all of the elements into various groups and periods according to their atomic numbers. This classification, known as the Periodic Table, reveals certain striking similarity and relationships between the elements in the same periods and in the same groups. Thus knowing, say, the physical and chemical characteristics of chlorine and then referring to the Periodic Table, we can predict accurately the physical and chemical characteristics of the other elements in that same group, namely fluorine, bromine and iodine. It is found that the elements in a certain group exhibit similar valence characteristics, as well as similar arrangement of electrons in their atomic structure. The periodic system has become an important tool of the research chemist.

Thus far we have taken up the elements oxygen, nitrogen and hydrogen, all of which are gases under ordinary conditions of temperature and pressure. Hydrogen, we find, is in Group Ia of the Periodic Table and has a valence of one. Nitrogen is in Group Vb of the Periodic Table and may exhibit any one of five different valences, namely 1, 2, 3, 4 or 5. Oxygen, we find, is in Group VIb of the Periodic Table and this element exhibits a valence of 2. It is too soon to try and draw any comparisons at this point, but as we take up other elements, we will soon begin to see the relationships which are brought out by the Periodic Table. You will find it interesting to refer to your chart of the Periodic Table as you study the various elements.

THE HALOGENS

The elements fluorine, chlorine, bromine and iodine are so similar in their properties and reactions that they form a natural family called the halogen group of elements. These non-metallic elements occur in Group VIIb of the Periodic Table and all of them exhibit a valence of one.

Fluorine is a pale yellow gas, exceedingly active and very poisonous. Chlorine is a greenish yellow gas, very soluble in water, heavier than air and also poisonous. Bromine is a dark reddish liquid, quite destructive to tissue. Iodine is a dark gray, lustrous solid. It is almost insoluble in water, but it is quite soluble in alcohol.

The halogens are so active chemically that they are never found free in nature. They are always combined with some other element or elements. The word "halogen" means "salt former". In combination with other elements, the halogens are among the most important compounds, both industrially and in our everyday life. For example, our common table salt, sodium chloride, is a combination of sodium and chlorine.

FLUORINE

Fluorine is the most reactive of the halogens and indeed it is the most active element known. Few substances are so stable that they are not attacked and decomposed by fluorine. The minerals fluorspar and cryolite are the chief sources of fluorine compounds. The element fluorine itself is very difficult to prepare as no chemical reagent is sufficiently powerful to liberate it from its compounds. To prepare pure fluorine, we must resort to the electric current.

Fluorspar (calcium fluoride) is the source of the interesting compound

hydrofluoric acid, H_2F_2, which dissolves glass. It is used for frosting and etching glass and of course, must be kept in rubber, wax or lead bottles.

Cryolite is important in the aluminum industry. It is a complex sodium and aluminum fluoride.

The difficulties and dangers involved in preparing and handling fluorine greatly retarded the development of commercial applications. Recently, however, improved techniques have made it available on a large scale and it is now being used in the manufacture of various compounds which can not be obtained in other ways. The field of usefulness of fluorine will probably be greatly widened within the next few years.

CHLORINE

Chlorine is by far the most abundant and the most important of the halogens. It occurs widely distributed in the form of various chlorides, the most common of which is sodium chloride. Since most chlorides are readily soluble in water, it is not surprising that they have been washed into the ocean in large amounts and that our chief source of chlorides is the salt beds left by the evaporation in past ages of land-locked seas.

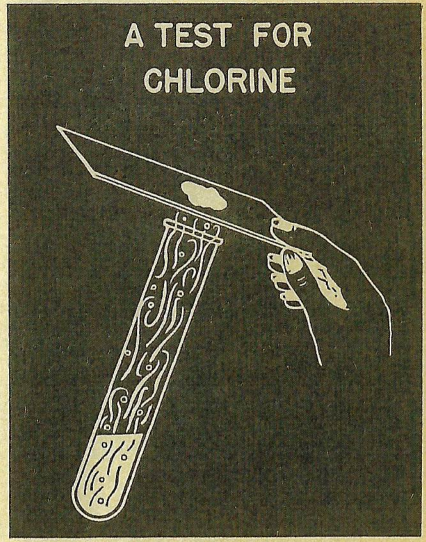
Chlorine gas is usually made from sodium chloride by an electrochemical process. The gas is liquefied by pressure and shipped in steel cylinders and tank cars for use in bleaching paper and textile fibers, in purifying water, and in many branches of chemical industry.

EXPERIMENT 148—How to Make Chlorine.

Place 1 measure of Bleaching Powder (No. 38) in a test tube, add about ½ inch of water, and shake vigorously, closing the mouth of the tube with your thumb. Now add 3 measures of Sodium Bisulfate (No. 7) to the white suspension in the tube. At once bubbles of gas begin to come off, and after a moment you can smell the characteristic odor of chlorine at the mouth of the tube. Be careful not to smell much of it, since it irritates the nasal passages.

EXPERIMENT 149—A Test for Chlorine.

Moisten the dull side of an exposed photographic negative and expose it to chlorine gas by holding it at the mouth of a test tube where chlorine is being generated. The minute particles of silver metal which form the dark image on the negative react with chlorine forming silver chloride. This is a white substance, and hence the dark portions of the negative take on a milky or cloudy appearance.



Bleaching powder is prepared by allowing chlorine gas to react with lime, calcium hydroxide, in large chambers. When mixed with water it forms calcium hypochlorite along with other compounds, and as you have just seen, chlorine is evolved when an acid is added to this mixture. Chlorine reacts slowly with water, liberating oxygen. At the moment when the oxygen atoms are first liberated, they are extremely active chemically, but after a small fraction of a second, they have all combined in pairs, forming oxygen molecules, O_2, in which state they pass off as gas. If we put colored materials into the mixture when oxygen is being set free, a few of the oxygen atoms will come in contact with the colored substance before they have time to pair up into molecules, and these very active oxygen atoms will react with, or oxidize, many colored materials, destroying the color. This is why bleaching powder is used for bleaching. Since it exerts a similarly destructive effect upon bacteria, bleaching powder is also a good disinfectant.

EXPERIMENT 150—Bleaching with Bleaching Powder.

Place 6 measures of Sodium Bisulfate (No. 7) and 4 measures of Bleaching Powder (No. 38) in a clean glass and add a few drops of water. Now obtain various scraps of paper, cloth, flowers, leaves, and a piece of litmus paper. Moisten these slightly and place them in the glass with the bleaching powder and sodium bisulfate. Cover the top of the glass with a saucer and let it stand for about an hour.

Some of the colors will be completely removed, others will be paler or changed to a different color, while some may not be affected at all. Printing ink and lead pencil marks owe their color to carbon and the bleaching powder has no effect upon them.

EXPERIMENT 151—Making Sodium Hypochlorite Bleach.

Fill a china cup about ¾ full of water and dissolve in it all of the sodium chloride (common table salt) that will go into solution after thoroughly stirring but without heating the solution. Take the small carbon rods from 2 old flashlight batteries and connect them, by means of wires wrapped around the rods or soldered to the brass caps, to 3 dry cells connected in series. Dip these carbons into the salt solution to serve as electrodes. Allow the current to pass for 20 minutes, stirring the solution from time to time so that the chlorine gas which is formed at the positive electrode during the electrolysis will react with the sodium hydroxide, which is produced around the negative electrode. This forms sodium hypochlorite. After about 20 minutes of operation, test the solution for bleaching properties by dipping into it a strip of Litmus Paper (No. 18).

EXPERIMENT 152—Another Method of Making Sodium Hypochlorite.

Fill a test tube about ⅓ full of water and add 4 measures of Bleaching Powder (No. 38). Shake the mixture thoroughly. In another test tube ⅓ full of water dissolve 4 measures of Sodium Carbonate (No. 4). Now pour the sodium carbonate solution into the test tube containing the bleaching powder and water. Shake this mixture for a moment and then filter. In this experiment you precipitated calcium carbonate and you now have in solution sodium hydochlorite and sodium chloride.

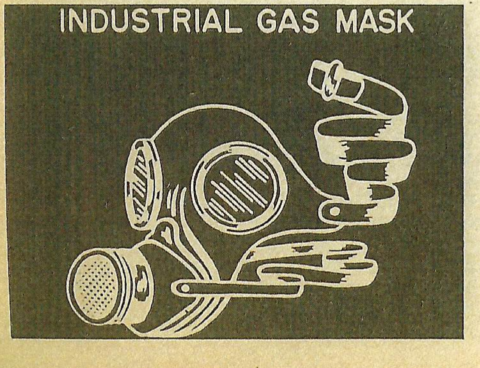
EXPERIMENT 153—Reaction of Chlorine with Sodium Thiosulfate.

Dissolve 2 measures of Sodium Thiosulfate (No. 16) in a test tube ¼ full of water. In another test tube place 2 measures of Bleaching Powder (No. 38), add ½ inch of water, and shake the tube vigorously. Then add 3 measures of Sodium Bisulfate (No. 7) and quickly fit the tube with the gas delivery tube and stopper. Pass the chlorine gas which is being evolved through the delivery tube into the sodium thiosulfate solution for a minute or two.

Sodium thiosulfate converts chlorine into sodium chloride, and while doing so it is itself changed into other substances. Among other things, free sulfur is formed, and that is the white precipitate which you see.

EXPERIMENT 154—A Gas Mask to Protect Against Chlorine.

Since sodium thiosulfate converts chlorine into a harmless compound it is an excellent protection against this irritating and poisonous gas. Chlorine was the first poison gas used in World War I and the gas masks used to protect against it contained sodium thiosulfate. Later, more deadly gases were introduced, and other chemicals had to be found to combat them.


INDUSTRIAL GAS MASK

Put 2 measures of Sodium Thiosulfate (No. 16) and 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water and shake it to dissolve the solids. Obtain a piece of cotton or linen cloth about 3 inches square and soak it in this solution for several minutes. When the cloth is thoroughly wet take it out and dry it carefully.

Now prepare chlorine from bleaching powder as in a previous experiment, and put the dry cloth over the mouth of the test tube while the chlorine is being generated. Smell at the mouth of the test tube next to the cloth and you will find that you cannot smell the chlorine gas. Put the cloth over your nose and try to smell the gas. You will find that the chemicals in the cloth react with the chlorine gas, changing it into a harmless compound.

Hydrochloric acid is one of the most important compounds of chlorine. A molecule of it is made up of one hydrogen atom and one chlorine atom, so its formula is HCl. It may also be called hydrogen chloride. Pure HCl is a gas, but it dissolves easily in water, and we usually use it in the form of a solution, called simply "hydrochloric acid". Such a solution has many uses — for cleaning steel preparatory to galvanizing or tinning, for making chlorides and numerous organic chemicals, in sugar refining, tanning, making glue, in dyeing, and in bleaching.

Hydrochloric acid is made commercially by heating sulfuric acid with common salt:

$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(sodium chloride)}}{2\,NaCl} \rightarrow \underset{\text{(sodium sulfate)}}{Na_2SO_4} + \underset{\text{(hydrochloric acid gas)}}{2\,HCl}$$

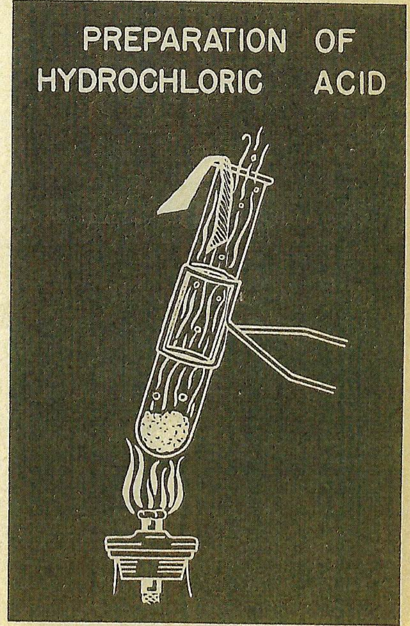
The gas which comes from the hot mixture is dissolved in water and ordinary hydrochloric acid solution is obtained. Very pure hydrochloric acid is often made by burning together hydrogen and chlorine gases:

$$H_2 + Cl_2 \rightarrow 2\,HCl$$

EXPERIMENT 155 —Preparation of Hydrochloric Acid.

In a dry test tube place 3 measures of Sodium Bisulfate (No. 7) and 1 measure of sodium chloride (common table salt). Heat this mixture strongly over a flame for a few moments, keeping your face away from the mouth of the tube. When fumes appear at the mouth of the tube, introduce a piece of moistened blue Litmus Paper (No. 18) and the hydrochloric acid gas will turn it red. Remove the test tube from the flame and smell cautiously at its mouth. You will notice the irritating odor of hydrochloric acid. Dissolving this gas in water would form the ordinary solution which we call hydrochloric acid.

Dip your glass rod into a bottle of household ammonia, and hold it near the mouth of the test tube in which hydrochloric acid is being generated. The white fumes are ammonium chloride, formed by the direct union of ammonia gas and hydrochloric acid gas:


PREPARATION OF HYDROCHLORIC ACID

$$\underset{\text{(ammonia)}}{NH_3} + \underset{\text{(hydrochloric acid)}}{HCl} \rightarrow \underset{\text{(ammonium chloride)}}{NH_4Cl}$$

BROMINE

Bromine, the next element of the halogen family, is a dark red liquid with a suffocating odor. When heated to about 60° Centigrade, it boils forming a choking red gas with an odor somewhat like that of chlorine. The vapors attack the eyes, nose, and throat, and the liquid can cause painful burns on the skin. Just as chlorine is less active than fluorine, so bromine is less active than chlorine, but it can form many useful compounds, both organic and inorganic.

Much bromine is obtained as a by-product from salt works. A plant on the coast of Texas extracts it directly from the sea. A ton of sea water contains only about two ounces of bromine (in the form of bromides), and 100,000 gallons of sea water must be pumped through the plant every minute to obtain a daily production of 45,000 pounds of bromine.

Bromine compounds are used in connection with anti-knock compounds in gasoline, in sensitizing photographic films, in tear-gas bombs, and in medicines.

IODINE

The last of the halogens is iodine, a grayish, lustrous solid, which is converted into a dark liquid and then into a heavy violet vapor when heated. Some iodine is obtained as a by-product in the refining of sodium nitrate (Chile saltpeter) in Chile, and also from the ashes of kelp, a kind of seaweed. Recently most of our iodine has been obtained by treatment of oil well brines, which in some localities contain from 30 to 70 parts per million of iodine in the form of sodium iodide.

Iodine is the least active of the halogens, and combines less readily with other elements. For this reason the other halogens are able to displace iodine from many of its compounds.

Small amounts of iodine in the diet are essential to human health. The thyroid gland in the neck uses it in making the complicated organic compound, thyroxin, $C_{15}H_{11}ONI_4$, which circulating in the blood, controls the rate of heat production in the tissues.

Iodine is an excellent antiseptic, being used chiefly in the form of a solution in alcohol, tincture of iodine. The compound, iodoform, CHI_3, is also a useful antiseptic. Large quantities of iodine in the form of silver iodide are used in the photographic industries, and iodine finds uses also in the dyestuffs industry and in chemical analysis.

EXPERIMENT 156—The Starch Test for Iodine.

Cut a thin slice from a raw Irish potato. Secure some tincture of iodine and allow 2 or 3 drops to fall on the slice of potato. Notice the beautiful blue coloration produced.

Starch (when cold) always gives this characteristic blue coloration in the presence of iodine.

EXPERIMENT 157—The Behavior of Starch Iodine Solution.

Crush a small piece of potato (¼ inch square), place it in a test tube and add ½ test tube of water. To this add a few drops of tincture of iodine as in the preceding experiment. Add enough to produce a beautiful blue color.

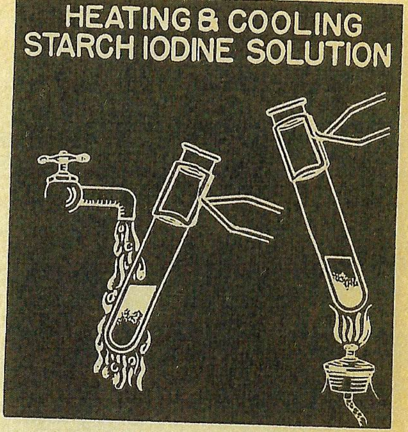
Warm the test tube gently over a flame, and watch the color gradually fade. Before the liquid begins to boil, it will lose all of its color. Hold the test tube in

cold water, and the blue color will slowly reappear. This experiment shows that solutions, to show the starch iodine color must be cold, for a hot solution will give no blue color.

Pour half of your blue solution into another test tube, and to one of the portions add ½ measure of Bleaching Powder (No. 38). Notice how rapidly the blue color is destroyed by the chlorine liberated from the bleaching powder.

To the other portion of the blue solution, add 4 measures of Sodium Carbonate (No. 4) and shake the tube. When a blue starch iodine solution is made alkaline, the free iodine is converted into sodium iodide and sodium hypoiodite and the blue color disappears. By neutralizing this solution with an acid, the blue color can be made to return because free iodine is again formed.

In testing for iodine or starch by means of the blue starch iodine solution, it is important to remember that you will not obtain the blue coloration: (1) if the solution is warm, (2) if the solution contains free chlorine, (3) if the solution is alkaline.



HEATING & COOLING STARCH IODINE SOLUTION

In some parts of the country the soil contains so little iodine that drinking water and vegetables do not supply enough iodine to the inhabitants, who therefore may suffer from goiter. For this reason the manufacturers of table salt add a very small amount of potassium iodide to the salt and call the product "iodized salt". Although there are only about 3 or 4 ounces of iodine in a ton of the salt, the starch test is easily able to detect it.

EXPERIMENT 158—Showing that Iodized Salt Contains Iodine.

Add about 2 teaspoonfuls of iodized salt and 5 measures of Sodium Bisulfate (No. 7) to a glass half full of water and stir until they are dissolved. Now prepare a starch solution by heating 1 measure of household starch in a test tube half full of water. As soon as the solution has started to boil, stop heating and allow it to cool. Add some of this starch solution to your salt solution. Then add a few tiny particles of Bleaching Powder (No. 38) and stir the liquid thoroughly. If a blue color, showing the presence of free iodine, does not appear, add another tiny portion of bleaching powder and stir vigorously. Continue adding bleaching powder until the liquid turns blue, but be very careful not to add too much at a time, because as you recall from a previous experiment, any excess of bleaching powder will destroy the blue color of the starch-iodine mixture. If you should add too much bleaching powder, the addition of another spoonful of iodized salt, along with, perhaps, a little more sodium bisulfate, will probably restore the blue color.

In one of our experiments with chlorine, we found that chlorine gas reacted with sodium thiosulfate. Since iodine is very similar to chlorine in chemical behavior, it is not surprising that it too reacts with sodium thiosulfate.

EXPERIMENT 159—Reaction of Iodine with Sodium Thiosulfate.

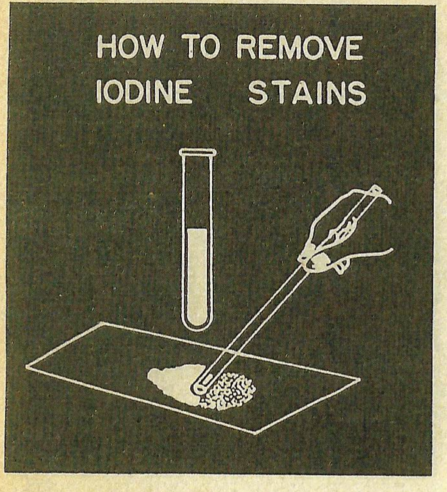
To a portion of the blue starch-iodine solution obtained in the preceding experiment, add 1 measure of Sodium Thiosulfate (No. 16) and shake it for a moment. The brown color of the water solution fades at once.

Instead of using the iodine solution left from the last experiment, you can make some more by adding 2 drops of tincture of iodine to a test tube half full of water. When the mixture has taken on a good brown color, drop in 1 measure of Sodium Thiosulfate (No. 16) and the color will disappear immediately.

The brown solution of iodine reacting with the thiosulfate forms two colorless compounds:

$$\underset{\text{(iodine)}}{I_2} + \underset{\text{(sodium thiosulfate)}}{2\,Na_2S_2O_3} \rightarrow \underset{\text{(sodium iodide)}}{2\,NaI} + \underset{\text{(sodium tetrathionate)}}{Na_2S_4O_6}$$

EXPERIMENT 160—How to Remove Iodine Stains.



This reaction is a good one to use for removing iodine stains from the skin or from fabric.

Make a solution of 4 measures of Sodium Thiosulfate (No. 16) in a test tube half full of water, and moisten the iodine stain with this solution. If the stain is not entirely bleached after standing a few minutes, repeat the application of thiosulfate solution.

Iodine, like chlorine, can enter into combination in the molecules of organic compounds. For example, just as we have methyl chloride, CH_3Cl, we have methyl iodide, CH_3I, and corresponding to chloroform (trichloromethane), $CHCl_3$, we have iodoform (triiodomethane), CHI_3. This substance is used as an antiseptic. It is very simply made by warming alcohol with iodine in the presence of an alkali.

EXPERIMENT 161—Making Iodoform.

Add 2 measures of Sodium Carbonate (No. 4) to a test tube half full of water. Shake the tube to dissolve it and then heat the solution to boiling. Now add to the hot solution about 15 drops of tincture of iodine, which of course is simply a solution of iodine in alcohol. Almost at once the brown color of the iodine disappears, and the liquid becomes cloudy due to the fine particles of iodoform which have formed. Let the test tube stand for a few minutes and the particles will grow larger and settle to the bottom. Iodoform is a yellow solid with a peculiar odor which may remind you of a doctor's office. The iodoform can be separated from the liquid by filtering.

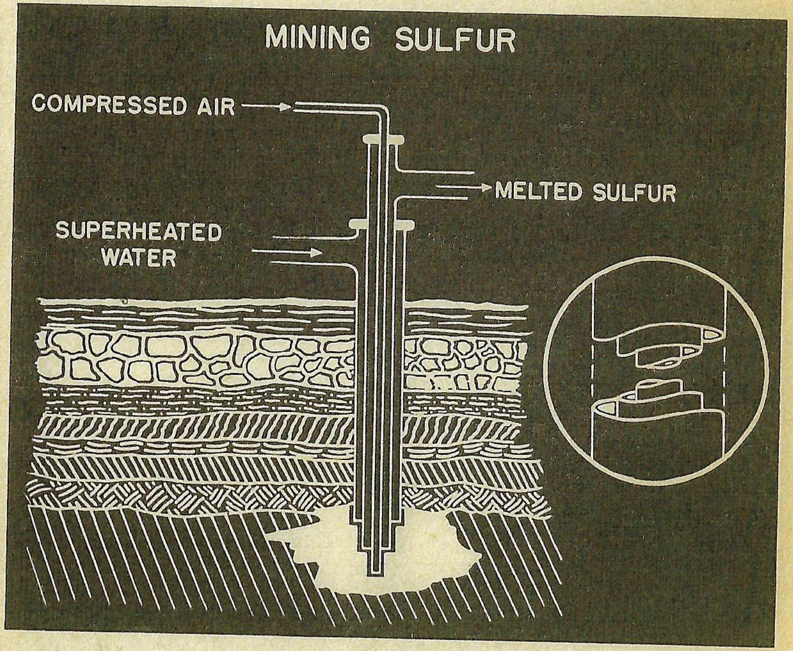
SULFUR

Sulfur is one of the world's most useful elements. You rarely see it in everyday life, but it enters directly or indirectly, into the manufacture of almost everything you use. Rubber, paper, metal, concrete, pigments, matches, medicines, and glass have all come into contact with sulfur or sulfur compounds at some stage of their manufacture. As you read the following pages, you will find many other examples of the part played by sulfur in the industries of the world.

Almost all of our sulfur comes from Texas and Louisiana, where it occurs in great deposits deep in the ground. For many years these deposits defied the efforts of mining engineers to work them in commercial quantities. Many attempts were made to obtain sulfur by digging deep shafts applying the methods used in coal mining. All of these attempts were hopeless failures because of quicksands, water, and poisonous gases.

It was not until 1903 that a successful method of tapping this great store of sulfur was put into operation. The credit for this achievement goes to Herman Frasch, an engineer who had long experience drilling wells in oil fields. He came to the conclusion after studying the problem that ordinary mining methods were out of the question, but that if the sulfur could be melted underground and pumped to the surface in a liquid condition, the problem would be solved.

As finally worked out, the Frasch process uses three pipes, one inside the other, driven down five hundred feet to the beds of sulfur. Through the outside pipe, about 8 inches in diameter, hot water is pumped down to the sulfur. This water is superheated, that is, it is heated under pressure to a temperature well above its boiling point. Coming into contact with the sulfur, this hot water melts it and forms a pool of molten sulfur at the bottom of the pipe. Hot compressed air is now forced down through the smallest pipe and, as it blows through the melted sulfur, a froth or foam of sulfur and hot water is formed. This is forced by the pressure of the hot water and the compressed air to flow up through the middle pipe to the surface. The melted sulfur is piped to large bins, where it is allowed to solidify. About 2,000,000 tons of sulfur are brought to the surface every year by this method. It is estimated that there are more than 60,000,000 tons of sulfur under the Gulf Coast of Louisiana and Texas.



Sulfur exhibits chemical properties quite similar to oxygen, which is what we would expect, since sulfur follows oxygen in Group VIb of the Periodic Table. Sulfur is a non-metallic element, and it commonly exhibits a valence of 2, 4, or 6.

Sulfur is a very versatile substance, assuming a great variety of forms. If we carefully heat some sulfur in a test tube, it melts easily, giving a pale yellow, oily liquid. Upon heating more strongly, the liquid becomes darker and thicker, and finally it is almost black and so thick that the test tube can be turned upside down without spilling it. Heated to a still higher temperature, it becomes again more fluid and at a high temperature can be made to boil.

EXPERIMENT 162—Liquid Sulfur.

Put 3 measures of Sulfur (No. 1) into a dry test tube and heat gently over a flame, keeping your face away from the mouth of the tube. Notice that the sulfur melts at once giving a straw-colored liquid. When you heat it more strongly, the color becomes darker, and the liquid thickens so much that it will not run out of the tube. Further heating causes it to boil, and the vapor condenses on the upper part of the test tube.

Just as the element carbon occurs in three forms, diamond, graphite and amorphous carbon, so the element sulfur may exist in any one of several forms, or allotropic modifications. Ordinary sulfur consists of crystals (sometimes very small) called *rhombic sulfur*. At a temperature above 96° C. these change, more or less slowly, into another kind of crystals, monoclinic sulfur, which when cooled changes back to the rhombic form.

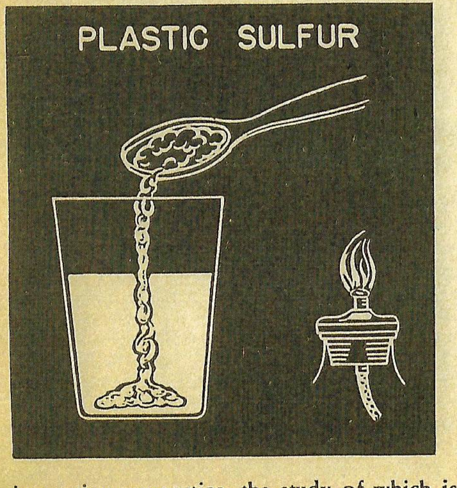
EXPERIMENT 163—Crystalline Forms of Sulfur.

Place 5 measures of Sulfur (No. 1) in your spoon and heat it gently over a candle flame until it is completely melted. Be very careful not to let it catch fire. Whenever fumes begin to come from it, remove it from the flame for a few moments. If it should start to burn, extinguish it by covering the spoon for a few seconds with a piece of metal or glass.

When the sulfur is all melted, remove the spoon from the flame and let it cool slowly. Soon the liquid will begin to solidify, or freeze. Watch closely, and when solidification is well started, tip the spoon so that the sulfur which is still liquid will run to one side, exposing the crystals already formed. Notice that these are shiny and generally somewhat elongated. They are monoclinic sulfur crystals.

Monoclinic sulfur is unstable at ordinary temperatures and tends to change over to the rhombic form, though the change may occur very slowly. Observe your monoclinic sulfur at intervals of half an hour, and you will soon see patches of dull yellow, which, after several hours, will spread over the whole of the material. Each of the monoclinic crystals has changed into a multitude of tiny rhombic crystals. Rubbing the monoclinic sulfur with a few grains of powdered sulfur generally helps to start the transformation.

By rapidly cooling melted sulfur, we can make an amorphous, or noncrystalline, variety having rather peculiar properties.



EXPERIMENT 164—Plastic Sulfur.

Melt 5 or 6 measures of Sulfur (No. 1) in your spoon by heating rapidly over a flame. If you wish, you can simply re-melt the sulfur remaining in the spoon from the preceding experiment. When completely melted, pour the sulfur into a glass of cold water. Remove the congealed sulfur and you will find that it is soft and almost rubbery. After a couple of days it will harden and change back to the ordinary form. This is amorphous sulfur.

When a substance is divided into extremely small particles, too small to be seen with an ordinary microscope, yet larger than single molecules, we say that it is in the colloidal state. Such particles when suspended in a liquid have a number of interesting properties, the study of which is a part of colloid chemistry.

EXPERIMENT 165—Collodial Sulfur.

Dissolve 1 measure of Sodium Thiosulfate (No. 16) in a test tube ¼ full of water. In another test tube ¼ full of water place 1½ measures of Sodium Bisulfate (No. 7) and shake it until it is completely dissolved. Now pour the sodium bisulfate solution in the second tube into the sodium thiosulfate solution in the first tube. Close the tube with your thumb and mix the two liquids by inverting it.

Watch the liquid for the next few minutes and soon you will see it begin to become cloudy. When you hold it up to the light and look through it, it will appear yellowish or brown, but if you hold it against a dark background and let light shine into it from the side, it will look blue. This is because the light passing into the liquid is scattered by the tiny sulfur particles, and more of the blue rays of light are deflected than of the others. The color of tobacco smoke has a similar origin.

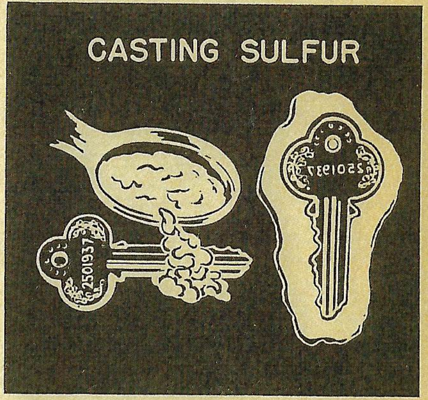
As the test tube stands, the particles grow larger, and presently the suspension will look nearly white unless the light passes directly through it, at which time it will look blue.

Particles as small as this pass right through the pores of ordinary filter paper. Try filtering some of your colloidal sulfur, and notice that the filtrate is about as cloudy as the original suspension.

Sulfur is sometimes used as a cement or as a casting medium, since it resists acids and is a good electrical insulator. It is poured while in a liquid form into a crevice or mold and allowed to solidify.

EXPERIMENT 166—Casting Sulfur.

Melt a small quantity of Sulfur (No. 1) in your spoon, being careful not to let it take fire. Pour the melted sulfur over a key or other small object. When it has solidified and the object is removed, you will find a perfect impression in the sulfur.



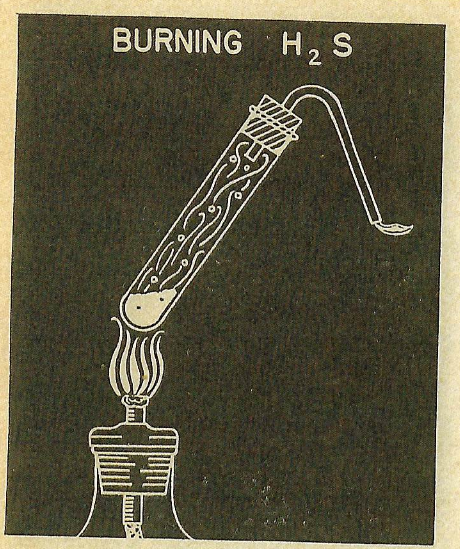
HYDROGEN SULFIDE

Hydrogen sulfide is a gas, each molecule of which is composed of two atoms of hydrogen and one atom of sulfur, H_2S. Everyone has smelled it, whether he realizes it or not, for it is the gas which is mainly responsible for the disagreeable odor of spoiled eggs and of the water in sulfur springs. Sewer gas also contains it. In spite of its foul odor, it is a very useful substance to the chemist, who uses it to precipitate the sulfides of the metals from solution in analytical work. The colors and characteristic reactions of the sulfides furnish a key to the metals which are present.

EXPERIMENT 167—Making Hydrogen Sulfide.

Cut from your candle a piece of paraffin about as big as the end of your little finger and place this together with 5 measures of Sulfur (No. 1) in a test tube. Heat the tube over a flame, and after a few minutes smell the hydrogen sulfide gas which is given off, first removing the test tube from the flame. Be careful not to smell very much of the gas, since it may give you a headache.

Paraffin is a hydrocarbon, that is, it is composed of hydrogen and carbon. When heated with sulfur, the hydrogen leaves the carbon and combines with the sulfur forming hydrogen sulfide.



EXPERIMENT 168—A Test for Hydrogen Sulfide.

Moisten a piece of Sulfide Test Paper (No. 31) and hold it in the mouth of the test tube in which hydrogen sulfide is being generated. The test paper will turn dark brown or black.

The test paper is impregnated with lead acetate, with which the hydrogen sulfide forms lead sulfide, a dark-colored, insoluble compound.

EXPERIMENT 169 — Burning Hydrogen Sulfide.

Fit the test tube in which hydrogen sulfide is being generated with the gas delivery tube and when the gas is coming off rapidly, apply a flame to the open end of the delivery tube, keeping your face at a little distance. The gas will burn with a blue flame.

When hydrogen sulfide burns in the air, the sulfur and hydrogen unite with oxygen, forming sulfur dioxide and water.

$$\underset{\text{(hydrogen sulfide)}}{2\,H_2S} + \underset{\text{(oxygen)}}{3\,O_2} \rightarrow \underset{\text{(sulfur dioxide)}}{2\,SO_2} + \underset{\text{(water)}}{2\,H_2O}$$

If, however, not enough air is present, only the hydrogen burns, and the sulfur is set free. Hold a cold piece of glass in the blue flame for a moment and notice the deposit of sulfur which forms on it.

EXPERIMENT 170—Testing for Hydrogen Sulfide in Illuminating Gas.

Test your gas supply for hydrogen sulfide by letting the gas from a jet flow against moistened Sulfide Test Paper (No. 31). Gas made from coal contains hydrogen sulfide, but this is supposed to be removed before the gas reaches the consumer.

EXPERIMENT 171—Hydrogen Sulfide from Boiling Cabbage.

The steam arising from cooking cabbage contains small amounts of hydrogen sulfide, formed from sulfur compounds in the cabbage. Hold a moistened strip of Sulfide Test Paper (No. 31) in the steam from a pot of boiling cabbage and after a minute or two it will darken, showing the presence of hydrogen sulfide.

EXPERIMENT 172—Tarnishing Metals with Hydrogen Sulfide.

Prepare some hydrogen sulfide by heating paraffin with sulfur as you did in one of the preceding experiments. Moisten a bright silver coin with water and hold it against the mouth of the test tube. After a minute or less, you will find a black stain of silver sulfide on the coin. Minute amounts of hydrogen sulfide in the air and other sulfur compounds in foods are the cause of the tarnishing of household silverware.

If instead of a silver coin, you use a penny, you will obtain a black stain of copper sulfide.

EXPERIMENT 173—Reaction of Hydrogen Sulfide with Chlorine.

Generate hydrogen sulfide gas as above by heating with paraffin, and pass the gas through your gas delivery tube into a test tube ¼ full of water. Let the gas bubble through the water for a minute or two.

Now generate chlorine gas in another test tube by adding a few drops of water to a mixture of 2 measures of Bleaching Powder (No. 38) and 4 measures of Sodium Bisulfate (No. 7). Fit the delivery tube and stopper to this tube, and let the chlorine gas bubble through the hydrogen sulfide solution which you have just made. In a few moments the liquid will become cloudy and soon you will have a dense suspension of minute particles of sulfur—colloidal sulfur.

The chlorine takes the hydrogen away from the sulfur, forming hydrochloric acid, and the sulfur is left suspended in the liquid in the form of tiny solid particles.

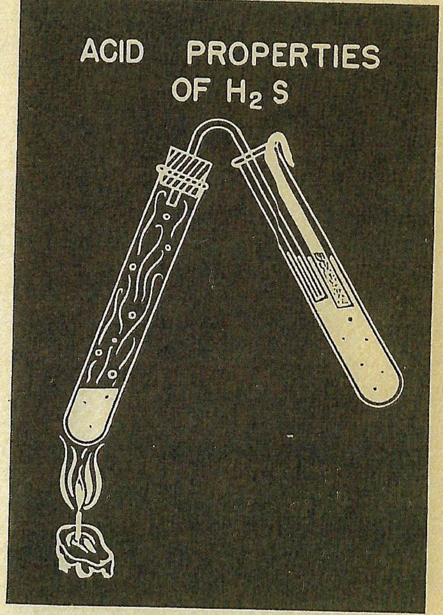
$$\underset{\text{(hydrogen sulfide)}}{H_2S} + \underset{\text{(chlorine)}}{Cl_2} \rightarrow \underset{\text{(hydrochloric acid)}}{2\,HCl} + \underset{\text{(sulfur)}}{S}$$

When dissolved in water, hydrogen sulfide forms a small number of hydrogen ions, and hence it is a weak acid.

EXPERIMENT 174—Acid Properties of Hydrogen Sulfide.

Fill a clean test tube ¼ full of water and drop in a small piece of blue Litmus Paper (No. 18). Now generate hydrogen sulfide gas as before by heating sulfur with paraffin and pass it through your gas delivery tube into the water containing the litmus paper. After the gas has bubbled through the water for a few moments, the litmus paper will turn red, showing that the solution of hydrogen sulfide is acid.

Being an acid, hydrogen sulfide forms salts with bases. These salts are called sulfides. Most of the sulfides are decomposed by strong acids with the liberation of hydrogen sulfide. This is sometimes a convenient way of preparing the gas.



EXPERIMENT 175—Decomposing a Sulfide.

Prepare some iron sulfide by heating 2 measures of Sulfur (No. 1) and 2 measures of powdered iron in your spoon until they combine. Place 1 measure of this compound in a test tube. Add 3 measures of Sodium Bisulfate (No. 7) and about ½ inch of water. Heat over a flame until the liquid begins to boil.

Notice the bubbles which immediately begin to rise from the particles of iron sulfide. After a few moments you will be able to smell the hydrogen sulfide at the mouth of the test tube. Moisten a piece of the Sulfide Test Paper (No. 31) and hold it at the mouth of the test tube. The brown discoloration on the paper proves that hydrogen sulfide is being evolved.

Just as oxygen reacts directly with many metals to form oxides, so sulfur reacts forming sulfides. Many sulfides can also be formed by passing hydrogen sulfide gas into a solution of some salt of the metal. Most sulfides are insoluble in water and hence precipitate out from the solution. Some of the sulfides dissolve readily in dilute acids and therefore will not form if the solution is acid. Others are not affected and may be precipitated even from an acid solution. These differences are used in analyzing mixtures of metals.

EXPERIMENT 176—Direct Formation of Silver Sulfide.

Place ½ measure of Sulfur (No. 1) on a bright silver coin and wrap it in several thicknesses of paper. After a few hours you will find a black spot of silver sulfide on the coin where the sulfur was in contact with it.

$$\underset{\text{(silver)}}{2\ Ag} \quad + \quad \underset{\text{(sulfur)}}{S} \quad \rightarrow \quad \underset{\text{(silver sulfide)}}{Ag_2S}$$

A silver coin is turned black in a few hours by a paste of mustard and water, showing that mustard contains sulfur compounds. Eggs also contain sulfur, which explains why silverware turns black when used with eggs.

EXPERIMENT 177—Direct Union of Copper and Sulfur.

Place ¼ measure of Sulfur (No. 1) on a copper coin. Hold the coin with your test tube holder and heat it until the sulfur melts. Let it cool, scrape off the sulfur, and examine the black film of copper sulfide.

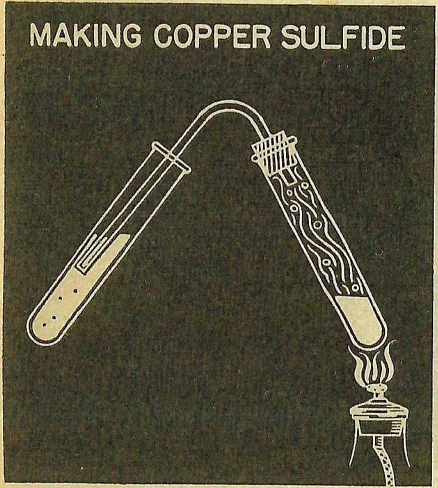
$$\underset{\text{(copper)}}{Cu} \quad + \quad S \quad \rightarrow \quad \underset{\text{(copper sulfide)}}{CuS}$$

EXPERIMENT 178—Proving That There is Sulfur in Rubber.

Rubber contains sulfur used in vulcanizing it. Wrap a rubber band around a silver coin, and after a few hours you will find black streaks of silver sulfide on it.

EXPERIMENT 179—Cobalt Sulfide.

Place 4 or 5 drops of Cobalt Chloride Solution (No. 27) and ½ measure of Sodium Carbonate (No. 4) in a test tube half full of water. A bluish precipitate of basic cobalt carbonate will form. Now pass hydrogen sulfide gas into the suspension and black cobalt sulfide will form from the particles of basic carbonate. The formula for cobalt sulfide is CoS.



EXPERIMENT 180—Copper Sulfide.

Place ¼ measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7) in a test tube ¼ full of water and warm the mixture gently until the solids have completely dissolved and the liquid is clear. This is a solution of copper sulfate.

Now generate hydrogen sulfide by putting a piece of paraffin or candle wax the size of the end of your little finger into a test tube along with 5 measures of Sulfur (No. 1). Fit the gas delivery tube and stopper to the test tube and heat the mixture. Pass the evolved hydrogen sulfide into the copper sulfate solution, and black copper sulfide will precipitate. A small quantity of sulfuric acid is also formed. This is one of the sulfides which does not dissolve easily in acids.

$$\underset{\text{(copper sulfate)}}{CuSO_4} \quad + \quad \underset{\text{(hydrogen sulfide)}}{H_2S} \quad \rightarrow \quad \underset{\text{(copper sulfide)}}{CuS} \quad + \quad \underset{\text{(sulfuric acid)}}{H_2SO_4}$$

EXPERIMENT 181—The Reaction of Hydrogen Sulfide with a Ferric Salt.

Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water. In another test tube, place a small piece of paraffin or candle wax with 5 measures of Sulfur (No. 1). Fit the gas delivery tube and stopper to this test tube and heat the mixture. As hydrogen sulfide gas is evolved, pass it into the test tube containing the ferric ammonium sulfate solution. Be very careful to keep the test tube in which hydrogen sulfide is being generated over the flame so that the ferric ammonium sulfate solution will not be sucked back into the hydrogen sulfide tube. After a short time, you will notice that a grayish white precipitate has formed in the tube which contained the ferric ammonium sulfate solution. This precipitate is sulfur formed by the oxidation of hydrogen sulfide by the ferric ions. The ferric ions are in turn reduced to ferrous ions. Were it not for this oxidation-reduction reaction in the case of the ferric salt, we would have obtained ferric sulfide, which is black.

EXPERIMENT 182—Zinc Sulfide

Place several small scraps of zinc metal in a test tube and add 3 measures of Sodium Bisulfate (No. 7) and ½ inch of water. As the zinc starts to dissolve, bubbles of hydrogen will come off. Heat the liquid to boiling to speed up the reaction, and let the test tube stand until no more bubbles come off. This may require 10 or 15 minutes. When the reaction is over, fill the test tube half full of water and bubble hydrogen sulfide gas through the solution. The white precipitate is zinc sulfide, ZnS.

Zinc sulfide will not form if any acid is present, and therefore, it is important to wait until all the acid has been used by reacting with the zinc metal.

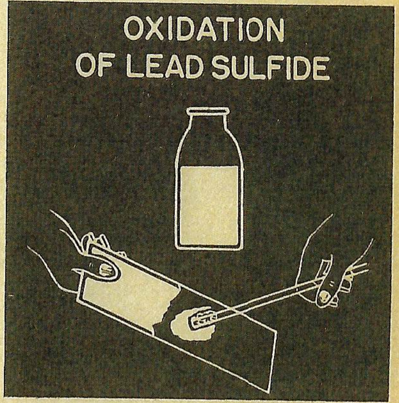
EXPERIMENT 183—Oxidation of Lead Sulfide.

Expose a moistened piece of Sulfide Test Paper (No. 31) to hydrogen sulfide gas. When it is thoroughly darkened, wet it with ordinary hydrogen peroxide, and it will turn white again. The hydrogen peroxide gives up oxygen to the black lead sulfide, converting it into white lead sulfate.

$$\underset{\text{(lead sulfide)}}{PbS} \quad + \quad \underset{\text{(hydrogen peroxide)}}{4\ H_2O_2} \quad \rightarrow \quad \underset{\text{(lead sulfate)}}{PbSO_4} \quad + \quad \underset{\text{(water)}}{4\ H_2O}$$

Many white paints contain lead carbonate, and traces of hydrogen sulfide in the air react with this compound to form lead sulfide, thus darkening the paint. Many famous paintings have suffered this fate, but fortunately the treatment with hydrogen peroxide restores their original appearance. For chemical laboratories, where hydrogen sulfide is common, zinc oxide is used as a white pigment in paint instead of the lead compound.

Many organic compounds contain sulfur in combination with carbon, hydrogen and other elements. Many of these have very unpleasant odors, as you will admit after performing the following experiment.



EXPERIMENT 184—Organic Sulfur Compounds.

Place 1 measure of Sulfur (No. 1) and 2 or 3 drops of glycerine in a dry test tube, and heat it slowly over a flame. As the glycerine begins to boil it will become reddish, and after a minute it will appear cloudy and have a red-brown color. Remove the test tube from the flame and let it cool slightly. At the mouth of the tube you will probably be able to recognize the odor of your friend hydrogen sulfide, tube you will probably be able to recognize the odor of your old friend hydrogen sulfide, but there will be another odor also, which may remind you of burning rubber. This is due to a mixture of organic sulfides and to a compound called allyl mercaptan.

SULFUR DIOXIDE

When sulfur burns in the air, the colorless gas, sulfur dioxide is produced. It has a sharp, choking odor, and is a heavy gas, being about twice as dense as air. Like chlorine, it is easily liquefied. Tremendous quantities are made by burning sulfur for use in the manufacture of sulfuric acid. Other important uses are for the preparation of sulfites for the paper industry, for bleaching various materials, and in numerous chemical manufacturing processes. Some of the household mechanical refrigerators use liquid sulfur dioxide in the same way as commercial plants use liquid ammonia.

EXPERIMENT 185—Making Sulfur Dioxide.

Heat one measure of Sulfur (No. 1) in your spoon over a flame until it melts and takes fire. Smell cautiously above the burning sulfur. Once you smell it, you will never forget the odor of sulfur dioxide.

$$S + O_2 \rightarrow SO_2$$

EXPERIMENT 186—Removing Fruit Stains with Sulfur Dioxide.

Stains caused by fruit juices are quickly bleached by sulfur dioxide. If you get some fruit stains on your hands, wet them with water and hold them over a little sulfur to which you have set fire. Fruit stains are also removed from fabrics by ths simple treatment. Wash your hands with soap and water as soon as the stain is bleached.

SULFURIC ACID AND THE SULFATES

In addition to sulfur dioxide, sulfur can form another oxide, containing an extra atom of oxygen. Its formula is SO_3, and its name is sulfur trioxide. When it reacts with water, it too forms an acid:

$$SO_3 + H_2O \rightarrow \underset{\text{(sulfuric acid)}}{H_2SO_4}$$

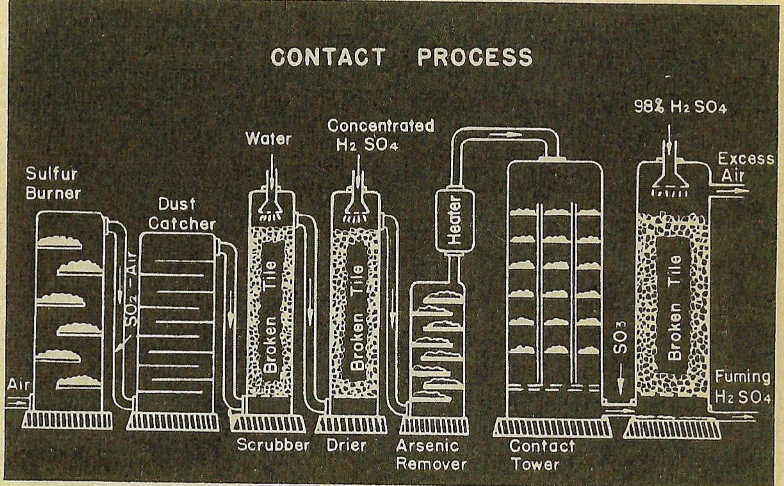
It is a general rule that when a non-metallic element forms more than one oxide, the one with the most oxygen in it gives the strongest acid, and therefore sulfuric acid is much stronger than sulfurous acid, which is formed when sulfur dioxide, SO_2, is dissolved in water.

Pure sulfuric acid is a heavy, colorless, oily liquid. It has a powerful attraction for water, and chars and destroys organic matter such as sugar, wood and paper. When diluted with water it readily dissolves many metals, giving off hydrogen and forming salts known as sulfates:

$$\underset{\text{(iron)}}{Fe} + \underset{\text{(sulfuric acid)}}{H_2SO_4} \rightarrow \underset{\text{(hydrogen)}}{H_2\uparrow} + \underset{\text{(ferrous sulfate)}}{FeSO_4}$$

Sulfuric acid is the cheapest of the strong acids and the most useful. There are no industries which do not use it directly or indirectly. More than a million tons a year are used in making fertilizers. The petroleum industry uses tremendous quantities in refining oils and gasoline, but the newer processes of solvent refining are beginning to drive it out of this field. The iron, steel, and other metallurgical industries would be seriously crippled without sulfuric acid. It is a necessity also in making a great variety of chemicals, in making explosives and various paints and pigments, and in the manufacture of rayon and cellophane. The liquid in the ordinary type of storage battery is moderately strong sulfuric acid.

There are two important methods for making sulfuric acid. In both of them, sulfur dioxide is the starting material. It is commonly made by burning sulfur, but sometimes it is recovered from waste smelter gases and sometimes it is made by roasting iron pyrites. In the chamber process, the sulfur dioxide is mixed in large lead chambers with oxygen and with oxides of nitrogen and steam. It is oxidized to sulfur trioxide, which combines with the steam forming sulfuric acid. Since the steam must be present in excess, the resulting acid is rather dilute.



When concentrated acid is desired, the contact process is generally used. Carefully purified sulfur dioxide is mixed with air and led over a catalyst. Originally, the catalyst was asbestos covered with finely divided platinum. The sulfur dioxide and oxygen combined on the surface of the tiny particles of platinum, forming sulfur trioxide. The platinum itself was not changed in the process and hence could be used over and over. Nowadays, a better catalyst, composed chiefly of vanadium pentoxide, is used.

$$2\ SO_2 + O_2 \text{ (catalyst) } \rightarrow 2\ SO_3$$

It is rather difficult to make sulfur trioxide combine directly with water to form sulfuric acid, and so the sulfur trioxide is absorbed first in previously made sulfuric acid, and then treated with water.

EXPERIMENT 187—Making Sulfuric Acid.

If you have a small amount of hydrogen peroxide, you can demonstrate the preparation of sulfuric acid. Obtain a half pint milk bottle or a similar wide mouth bottle and cut a small square of cardboard that you can place across the top to prevent the escape of gas. Now place 5 measures of Sulfur (No. 1) in your spoon, lay the spoon on a tin plate and light the sulfur. When the sulfur is burning well, hold the mouth of the bottle about an inch above the spoon so that the sulfur dioxide coming off from the burning sulfur will rise into the bottle. Hold the bottle in this position for several minutes until it is well filled with sulfur dioxide gas. Now quickly slip the piece of cardboard over the mouth of the bottle and turn it right side up, keeping the cardboard in place. Now add some hydrogen peroxide to the bottle (an amount about equal to a tablespoonful) and quickly replace the cardboard over the mouth of the bottle. Holding the cardboard in place with your fingers, shake the bottle around so that the hydrogen peroxide will come in contact with the sulfur dioxide gas. Shake the bottle for a moment or two, then remove the cardboard and test the liquid in the bottle with a piece of blue Litmus Paper (No. 18). You will find that the litmus paper turns a bright red, indicating that you have an acid.

$$\underset{\text{(sulfur dioxide)}}{SO_2} + \underset{\text{(hydrogen peroxide)}}{H_2O_2} \rightarrow \underset{\text{(sulfuric acid)}}{H_2SO_4}$$

In a clean test tube ¼ full of water dissolve 1 measure of Strontium Chloride (No. 10) and add the few drops of sulfuric acid which you have made. A white precipitate of strontium sulfate forms, proving that the acid which you made is actually sulfuric.

$$\underset{\text{(sulfuric acid)}}{H_2SO_4} + \underset{\text{(strontium chloride)}}{SrCl_2} \rightarrow \underset{\text{(strontium sulfate)}}{SrSO_4} + \underset{\text{(hydrochloric acid)}}{2\ HCl}$$

EXPERIMENT 188—Precipitating Strontium Sulfate.

In a test tube ¼ full of water, dissolve 1 measure of Aluminum Sulfate (No. 12) and in another test tube ¼ full of water dissolve 1 measure of Strontium Chloride (No. 10). Pour the 2 solutions together and you will obtain a heavy white precipitate of strontium sulfate. This is the same thing chemically as the mineral celestite. The precipitation of strontium sulfate is the same reaction you used in the preceding experiment to test your sulfuric acid.

EXPERIMENT 189—Testing for Sulfuric Acid in a Storage Battery.

By means of a glass tube, remove a drop of the electrolyte from an automobile storage battery and drop it into a test tube ¼ full of water. Test this solution with a piece of blue Litmus Paper (No. 18) and you will find that the paper is immediately turned red, showing that the liquid is strongly acid.

In another test tube ¼ full of water, dissolve 1 measure of Strontium Chloride (No. 10). Pour a few drops of the diluted electrolyte from the first tube into the strontium chloride solution. The white precipitate of strontium sulfate proves that the electrolyte contains sulfate ions and since we have already shown it to be acid, we conclude that sulfuric acid must be present. This experiment will not work with the electrolyte from an alkali storage cell which contains sodium hydroxide solution instead of sulfuric acid.

SODIUM THIOSULFATE

Sodium thiosulfate is a salt of the very unstable thiosulfuric acid, $H_2S_2O_3$. Notice that this acid differs from sulfuric in having one more sulfur atom and one less oxygen atom in its molecule. It loses the extra sulfur atom very readily.

EXPERIMENT 190—Separating Sulfur from Thiosulfuric Acid.

Dissolve 2 measures of Sodium Thiosulfate (No. 16) in a test tube half full of water and drop in 2 measures of Sodium Bisulfate (No. 7). Close the tube with your thumb and shake it well. Soon the solution will appear slightly turbid, and gradually a thick precipitate of sulfur will form. At the mouth of the tube you can smell sulfur dioxide.

Acids liberate thiosulfuric acid from sodium thiosulfate, and the unstable thiosulfuric acid then slowly decomposes into sulfurous acid and sulfur.

Although the salts of thiosulfuric acid, the thiosulfates, are much more stable than the free acid, they too break down when heated. If water is present, hydrogen sulfide is given off.

EXPERIMENT 191—Making Hydrogen Sulfide from Sodium Thiosulfate.

Place 1 measure of Sodium Thiosulfate (No. 16) in a dry test tube and heat it strongly over a flame. This salt contains water of hydration, part of which is first driven off, condensing in droplets on the upper part of the test tube. On continued heating, the remaining water will react with the thiosulfate, forming hydrogen sulfide. When the solid begins to turn dark, remove the test tube from the flame and smell carefully at its mouth for the familiar odor of hydrogen sulfide. Hold a moistened piece of Sulfide Test Paper (No. 31) at the mouth of the tube, and it will soon darken. Some free sulfur is also formed by the decomposition as you can see by examining the wall of the test tube just above the dark residue in the bottom.

EXPERIMENT 192—Formation and Decomposition of Ferric Thiosulfate.

Dissolve 2 measures of Sodium Thiosulfate (No. 16) in a test tube half full of water, and in another test tube ¼ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). Add a few drops of the ferric ammonium sulfate solution to the sodium thiosulfate solution, and you will obtain a dark red color. This is a solution of ferric thiosulfate. In a few seconds you will notice it beginning to fade, and after a minute or two the solution will be nearly colorless, since ferric thiosulfate decomposes rapidly forming ferrous tetrathionate, which has no color.

Since the thiosulfates, under certain conditions, may decompose producing hydrogen sulfide, the sulfides of some of the metals may be precipitated by means of sodium thiosulfate.

EXPERIMENT 193—Precipitating Cuprous Sulfide.

To a test tube ¼ full of water add ½ measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7). Warm over a flame until the solution becomes clear. It now contains copper sulfate. Heat it just to boiling and add 1 measure of Sodium Thiosulfate (No. 16). Boil for a few moments, taking care not to let the liquid boil over. First a yellowish color appears, then the solution becomes cloudy, and soon there is a thick brown precipitate. Let the tube stand for a few minutes, occasionally warming it again over the flame. The precipitate will slowly darken and finally a heavy black substance will settle to the bottom. This is cuprous sulfide, Cu_2S. The sulfide of copper which you precipitated with hydrogen sulfide in a previous experiment was cupric sulfide, CuS. The two sulfides resemble each other closely.

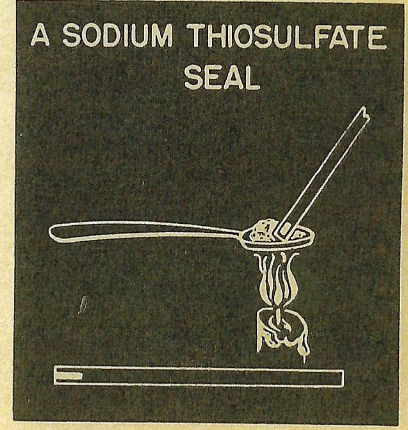
Sodium thiosulfate finds wide application in chemistry and in industry. The "hypo" which the photographer uses is nothing but sodium thiosulfate. It has the property of dissolving silver salts from negatives and prints. During the process of development, the silver salts which have been exposed to light are changed into finely divided metallic silver. After development, the films or prints are placed in the hypo "fixing" bath, which dissolves out the unchanged silver salts, but leaves behind the black deposit of metallic silver which forms the image. The fixing bath thus removes all the remaining light-sensitive compounds and makes the picture permanent.

In earlier sections we saw that sodium thiosulfate reacts with chlorine and iodine. This is the basis of its use as an "anti-chlor" to remove excess chlorine from textiles after bleaching. Its reaction with iodine is an important one in analytical chemistry. Other uses are in certain dyeing operations and in the sugar industry to prevent fermentation.

When heated, sodium thiosulfate melts easily in its water of crystallization and solidifies on cooling. Because of this behavior it has been used for sealing glass tubes containing explosives in underwater torpedoes.

EXPERIMENT 194 — A Sodium Thiosulfate Seal.

Melt 4 measures of Sodium Thiosulfate (No. 16) in your spoon by heating gently over a flame. Be careful not to heat it more strongly than necessary. Dip the end of a piece of glass tubing into the melt, tilt it a little and remove it in such a way that the end remains filled with the melted thiousulfate. When it cools you will find that the tube is tightly closed. Such seals, of course, are only of temporary value, since the solid salt is dissolved by water.


A SODIUM THIOSULFATE SEAL

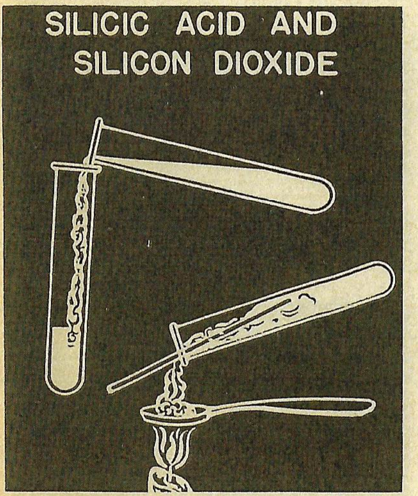
Silicon and Silicates

About a quarter of the earth's crust is silicon in combination with other elements. Quartz, sand, and sandstone are composed of more or less pure silicon dioxide, SiO_2, and many common minerals such as clay, feldspar, mica, soapstone, asbestos, and garnet consist of silicates of various metals. Free silicon is very similar to boron. It is not easy to prepare pure, but in the form of an alloy with iron it finds wide application in the iron and steel industry.

Silicon is a non-metallic element occuring in Group IVb of the Periodic Table. It exhibits a valence of 4.

The compounds of silicon are very numerous and some of them are of great importance. In addition to the natural compounds of silicon mentioned above, several artificial ones are worth mentioning. Silicon carbide, SiC, made from sand and coke in electric furnaces, is sold commercially as Carborundum for use in making grinding wheels. Sodium silicate is made by heating sand with sodium carbonate or sodium hydroxide. When dissolved in water it is called water glass and is used as an adhesive and in the manufacture of soaps, pottery glazes, and many other materials.

EXPERIMENT 195—Making Silicic Acid.



Put ½ inch of Sodium Silicate Solution (No. 19) in a test tube and add water until the tube is ¼ full. Shake the test tube to mix the liquids.

In another tube put 4 measures of Sodium Bisulfate (No. 7) and fill the tube ⅓ full of water. Shake the tube until the solid is completely dissolved.

Pour the sodium silicate solution into the sodium bisulfate solution and mix. A jelly-like precipitate will form, and in a few minutes all the liquid in the tube will become solid.

Save this silicic acid jelly for use in the next experiment.

EXPERIMENT 196—Making Silicon Dioxide.

Remove some of the silicic acid formed in the preceding experiment from the test tube by means of a stirring rod and put some of it in your spoon. Heat the spoon strongly in an alcohol lamp flame for 4 or 5 minutes, or until the precipitate is entirely dried and has been reduced to a white solid. This substance is silicon dioxide, the same thing chemically as pure sand or quartz, mixed with sodium sulfate, a by-product of the reaction in the preceding experiment.

EXPERIMENT 197—Soluble Glass (Sodium Silicate).

Paint a thin film of Sodium Silicate Solution (No. 19) on a sheet of paper and let it dry for 10 or 15 minutes. Notice the smooth, transparent, glass-like film which results.

Paste together two sheets of paper or two blocks of wood, using sodium silicate solution as the adhesive. You will find that it makes an exceptionally strong adhesive, and it is often used for this purpose.

Silicates of sodium and potassium are the only silicates which are readily soluble in water.

The most important artificial silicates are glass and portland cement. Glass is to be discussed in a later section. Portland cement is made by strongly heating a mixture of calcium carbonate (limestone) and a siliceous clay. The calcium unites chemically with the silica and alumina of the clay, producing a substance which when very finely ground and mixed with water, sets to a solid stone.

Clay is a combination of silica with aluminum oxide and water. It is the combined water which gives the property of plasticity, and makes it possible to mold it into pottery and bricks. When the molded article is "burned," that is, heated strongly, the combined water is driven off and the clay is made permanently hard.

Most metallic silicates are insoluble in water. They are formed as beautiful precipitates when a soluble silicate is added to a solution of the metallic salt.

EXPERIMENT 198—Aluminum Silicate.

Place 2 measures of Aluminum Sulfate (No. 12) in a test tube and fill the tube half full of water. shaking it to completely dissolve the solid. Now add 2 or 3 drops of Sodium Silicate Solution (No. 19) and notice the thick precipitate of aluminum silicate which is immediately formed.

EXPERIMENT 199—Cobalt Silicate.

Add 3 or 4 drops of Cobalt Chloride Solution (No. 27) to a test tube half full of water. Now add 2 or 3 drops of Sodium Silicate Solution (No. 19). In this case, a beautiful blue precipitate of cobalt silicate is formed. This is one of the most colorful of the entire series of silicates.

EXPERIMENT 200—Copper Silicate.

Place 3 measures of Sodium Bisulfate (No. 7) and 1 measure of Azurite (No. 39) in a test tube and fill the tube half full of water. Heat the liquid for a few moments until a clear blue solution of copper sulfate is obtained. Add to the copper sulfate solution 2 or 3 drops of Sodium Silicate Solution (No. 19), and examine the blue precipitate of copper silicate which is formed.

EXPERIMENT 201—Ferric Silicate.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water and add 2 or 3 drops of Sodium Silicate Solution (No. 19). A very pretty, reddish brown precipitate of ferric silicate will be formed.

EXPERIMENT 202—Magnesium Silicate.

Dissolve 2 measures of magnesium sulfate (Epsom salts) in a test tube half full of water. Add to this solution 2 or 3 drops of Sodium Silicate Solution (No. 19). A white precipitate of magnesium silicate will be formed.

EXPERIMENT 203—Strontium Silicate.

Dissolve 2 measures of Strontium Chloride (No. 10) in a test tube half full of water and add 2 or 3 drops of Sodium Silicate Solution (No. 19). A bulky, white precipitate will form and upon shaking the test tube the precipitate will fill the whole tube.

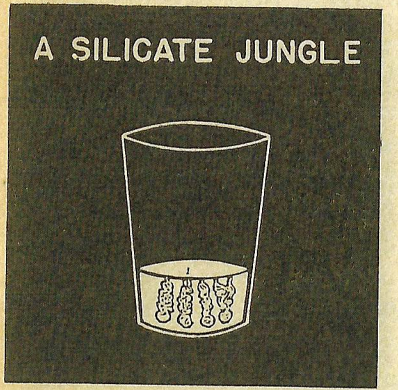
EXPERIMENT 204—Zinc Silicate.

Place a small piece of zinc metal and 2 measures of Sodium Bisulfate (No. 7) in a test tube. Fill the tube ¼ full of water, and warm the solution for a few moments until some of the zinc dissolves. Now hold the tube in a glass of cold water until it becomes cool. Add to the solution of zinc sulfate thus formed 2 or 3 drops of Sodium Silicate Solution (No. 19). A thick precipitate of zinc silicate will be formed.

EXPERIMENT 205—Reaction of Ammonium Chloride with Sodium Silicate.

We have been making silicates by adding sodium silicate solution to solutions of other salts, and you might ask now whether we could make ammonium silicate by adding sodium silicate solution to an ammonium salt, for example, ammonium chloride. The answer is that if ammonium silicate is formed, it decomposes at once, since its acid, silicic acid and its base, ammonium hydroxide, are both so weak that they cannot hold each other. Furthermore, ammonium hydroxide decomposes very easily into ammonia gas and water. Another way of stating the matter is to say that ammonium silicate is completely hydrolyzed by water, forming silicic acid and ammonium hydroxide. The ammonium hydroxide then proceeds to decompose.

Dissolve 2 measures of Ammonium Chloride (No. 9) in a test tube ¼ full of water and add 2 or 3 drops of Sodium Silicate Solution (No. 19). Close the mouth of the tube with your thumb at once and shake vigorously. The contents of the tube turn to a white slush or jelly of silicic acid, and if you smell at the mouth of the tube you can detect the ammonia formed in the reaction.



EXPERIMENT 206 — A Silicate Jungle.

Place 2 spoonfuls of Sodium Silicate Solution (No. 19) in a tumbler and fill it ¼ full of water. Mix the two liquids and drop in crystals of some of the following salts: Ferric Ammonium Sulfate (No. 21), Aluminum Sulfate (No. 12), Strontium Chloride (No. 10) and magnesium sulfate (Epsom salts). Let the glass stand quietly for a few hours and observe the plant-like growth which springs up from the crystals. This is formed of silicate membranes.

Two Other Non-Metallic Elements

There are two other non-metallic elements which should be mentioned at this point. They are boron and phosphorus.

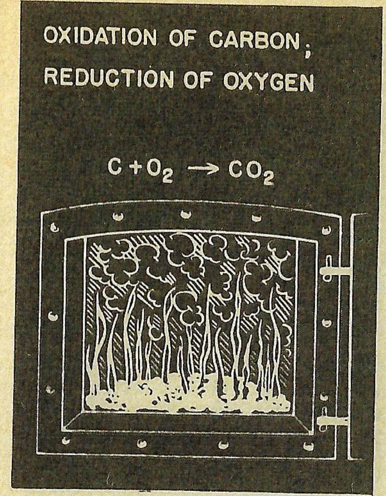
Boron occurs in Group IIIa of the Periodic Table and exhibits a valence of 3. The most important compound of boron is sodium borate or borax, $Na_2B_4O_7.10H_2O$. The element boron is exceedingly hard, being second only to the diamond.

Phosphorus occurs in Group Vb of the Periodic Table. Like nitrogen, which it follows in Group Vb, phosphorus exhibits five states of valence, from 1 to 5. Phosphorus precedes sulfur in period 3 of the Periodic Table and like sulfur, occurs in two physical forms, namely, white phosphorus and red phosphorus.

Phosphorus is a very interesting element and forms a large number of useful compounds, both with metals and with other non-metals. Phosphorus is found in nature in the form of phosphates, the most abundant being the mineral phosphorite, a calcium phosphate, $Ca_3(PO_4)_2$. Calcium phosphate is the basis of most of our compounds of phosphorus, including superphosphate, a calcium hydrogen phosphate, $Ca(H_2PO_4)_2.H_2O$, which is an important ingredient of fertilizer. Trisodium phosphate, known as TSP, is used extensively in soap powders and as an aid in cleansing.

Oxidation and Reduction

Many of the most important chemical processes of industry and of daily life depend upon reactions in which the valence of elements changes. Such reactions are called oxidation-reduction reactions. The burning of fuels,



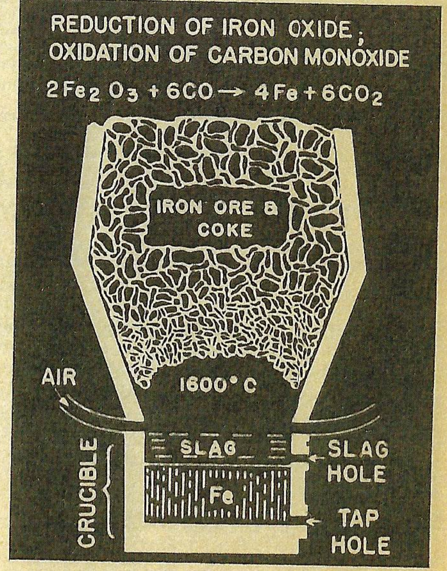
the smelting of metals, the operation of dry-cells and of storage batteries, the rusting and tarnishing of metals, and the growth of plants and animals are intimately connected with oxidation-reduction reactions. When iron is not combined with any other element, you can say that its valence is zero because it is not using any of its capacity for combining. Suppose now that one allows the iron to react with oxygen, forming ferrous oxide, FeO:

$$2\,Fe + O_2 \rightarrow 2\,FeO$$

The valence of iron in this compound is plus two, so the iron has increased two in valence. An increase in valence is called oxidation. Now let's see what has happened to the oxygen. The free oxygen, before the reaction, had a valence of zero because it was not combined with another element. After the reaction its valence is minus two. Its valence then went from zero to two below zero, that is, it decreased two. A decrease in valence is called reduction. In this reaction the iron was oxidized by the oxygen, and the oxygen was reduced by the iron. Whenever an atom is oxidized, some other atom must always be reduced.

We have already met examples of oxidation and reduction reactions. When we burned sulfur, we oxidized it, and simultaneously an equivalent amount of oxygen from the air was reduced by combining with the sulfur.

When we heated iron with sulfur, making iron sulfide, the reaction was very similar to the burning of iron in air. The only difference was that the iron combined with sulfur instead of with oxygen. Chemists extend the word oxidation to include the combination of non-metals with metals, regardless of whether there is any oxygen in the reaction or not. A similar statement goes for reduction also.



Now there is still more to this matter of oxidation and reduction. We find that many of the elements can combine with various fixed amounts of other elements. An atom of iron, for example, can combine with two atoms of chlorine, or with three atoms. We say, therefore, that it may have a combining capacity of either two or three (these numbers are considered as plus, or positive, since iron forms positive ions), while chlorine has a combining capacity of minus one.

When we liberate iodine from sodium iodide solution by passing in chlorine:

$$2\,NaI \quad + \quad Cl_2 \quad \rightarrow \quad 2\,NaCl \quad + \quad I_2$$

we change its combining power from minus one to zero. This is, in a certain sense, an increase in combining power, and hence the iodine is oxidized. The chlorine at the same time changed to sodium chloride, in which its combining power is minus one. Its change, then, is from zero (free chlorine) to minus one, and thus it is reduced. Oxidation and reduction always go together.

All branches of chemistry must deal with oxidation and reduction reactions—metallurgy is based upon them, electrochemistry and phyisical chemistry make use of them, organic chemists continually discuss them, and they lie at the very root of bio-chemistry, the chemistry of life.

EXPERIMENT 207—Formation of Iron Oxide.

Hold a small bright iron nail with a pair of pliers and heat it in the flame of the alcohol lamp for a few minutes. Plunge it into a glass of water and when it is cooled remove it and notice the changed appearance. The black color is due to iron oxide.

EXPERIMENT 208—Formation of Copper Oxide.

Hold a piece of bare copper wire with a pair of pliers and heat it in the flame of the alcohol lamp. The copper will impart a beautiful green coloration to the flame and when the copper is removed and plunged into a tumbler of water a black scaly deposit of copper oxide will form.

EXPERIMENT 209—Formation of Zinc Oxide.

Cut off a piece of zinc metal about the size of your fingernail and scratch it with a knife until it is shiny. Drop this into a test tube and heat it strongly over the alcohol lamp flame. After a few minutes you will notice a white deposit of zinc oxide appearing on the metal.

An interesting example of the oxidation of metals is the process of cutting with the oxy-acetylene flame. Oxygen and acetylene gases are mixed and the gas is lighted. This makes a flame so hot that it may be used in cutting through iron and steel of almost any thickness. The metal is heated to its kindling temperature and burned away.

Oxidation is not always accompanied by light or by noticeable heat. For example, when iron rusts it combines slowly with oxygen. When wood decays the products are nearly the same as those formed when it burns. Just as much heat is given off when iron rusts as when it is burned at a high temperature, but in the rusting the action requires such a very long time and the heat is given off so slowly that it is not noticeable, as it has no chance to accumulate.

EXPERIMENT 210—The Slow Oxidation or Rusting of Iron.

Put 2 or 3 measures of powdered iron in a tumbler and moisten it with a very few drops of water. Set the tumbler aside for a few hours and then notice the rust which is beginning to form. The iron rust which is formed in this experiment is very similar to the iron oxide which is formed when iron is burned at a high temperature.

An excellent example of reduction is the action of thermit, a mixture of powdered aluminum metal and iron oxide. Here the aluminum metal is used to combine with oxygen from the iron oxide just as carbon is used in the blast furnace.

In making welds, thermit is placed in a crucible having a small tap hole in the bottom and the work to be welded is contained in a mold immediately below the tap hole. The thermit is ignited and a violent reaction immediately takes place. The aluminum metal takes the oxygen away from the iron oxide forming aluminum oxide and metallic iron.

$$\underset{\text{(aluminum)}}{2\,Al} \quad + \quad \underset{\text{(iron oxide)}}{Fe_2O_3} \quad \rightarrow \quad \underset{\text{(aluminum oxide)}}{Al_2O_3} \quad + \quad \underset{\text{(iron)}}{2\,Fe}$$

The heating resulting from the union of the aluminum and the oxygen is intense enough to melt the iron. The tap hole in the bottom of the crucible is then opened and the molten iron runs into the mold, welding the piece which is held there.

In a broader sense the terms "oxidation" and "reduction" refer to reactions which are equivalent to the addition or removal of oxygen, even if no oxygen is actually present.

Metals

Of all the chemical elements, more than three-fourths are classed as metallic,which means that they can be made to combine with oxygen forming oxides which have basic properties and which unite wit hacids to form salts.

The most striking physical characteristics of metals are their luster, their ability to conduct electric current and to conduct heat. Everyone is familiar with the common metals, such as iron, copper, lead and tin and with their characteristics of weight, strength and metallic luster. There are, however, some metals which lack most of these properties and the real distinction between metals and non-metals is based upon chemical propreties.

Very few metals are found in nature in the pure state. Mostly they occur as ores which consist of various compounds of the metals. Iron, for example, occurs in the form of iron oxide, iron sulfide and iron carbonate. Copper occurs mostly as copper sulfide and copper carbonate. Lead occurs principally as lead sulfide. Gold, on the other hand, occurs in the free state mixed with various other minerals.

In order to obtain pure metals from the ores in which they occur in nature, it is necessary to devise certain methods of liberating the metals from their combinations. This is done mostly through oxidation and reduction reactions. In many cases the ore is first roasted or oxidized to remove impurities. The resulting oxide of the metal is then subjected to a process of reduction and the metal is obtained in the free state. Further operations may then be required to purify it. The art and science of extracting metals from their ores is called metallurgy.

ALLOYS

In many cases molten metals dissolve in one another freely. The mixtures, or compounds, resulting are called alloys. Many alloys have entirely different properties from those of the metals which are used in making them.

An amalgam is an alloy of mercury with some other metal. Copper forms alloys with many other metals. Brass contains various properties of copper and zinc. A yellow brass, for example, may run about 30% zinc and 70% copper. A typical bronze contains about 5% of tin, 15% of zinc, a small quantity of lead and about 75% of copper. Gunmetal and bell metal are both mixtures of copper and tin. German silver is a mixture of copper, zinc and nickel.

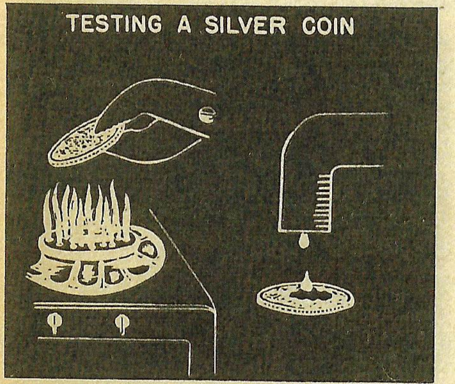
Pure iron is a relatively soft metal and could not be employed for the uses that are demanded of wrought iron and steel were it not mixed with other materials to form tough, durable alloys called steel.

Most coins are alloys of various metals, as the pure metal is in most cases too soft to stand the wear to which they are subjected. Our 5 cent piece, for example, is an alloy containing 75 per cent copper and 25 per cent nickel. Our silver coins contain 90 per cent of silver and 10 per cent of copper.

EXPERIMENT 211—Testing a Penny for Copper.

Heat a copper cent in a gas flame or over a coal fire until it becomes red hot. Notice the very pretty colors which play over the surface of the coin while you are heating it. Now let a few drops of water fall on the coin and when it becomes cool notice the formation of black copper oxide which covers it.

EXPERIMENT 212—Testing a Silver Coin to Show the Presence of Copper.



Repeat the preceding experiment with a silver coin, and notice that a somewhat lighter coat of black copper oxide is obtained in this case. Pure silver will not oxidize in the air even when heated. The black coating is copper oxide, and the presence of this oxide on the silver proves that coins are not pure silver.

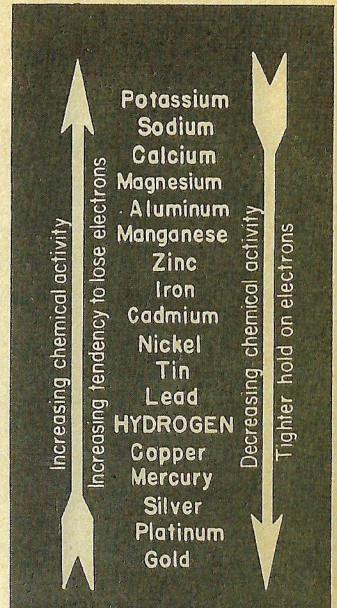
EXPERIMENT 213—Testing a Nickel Coin to Show that It Contains Copper.

Heat a 5-cent piece in a gas flame or over a coal fire until it becomes red hot, and then let a few drops of water fall on it. You will notice the same formation of black copper oxide which was obtained when you heated the cent.

The melting point of an alloy is often lower than the melting point of either of the metals of which it is composed. Alloys with low melting points are frequently used in automatic sprinkler systems and on fire doors so that if a fire starts and the metal gets hot it melts and releases the water from the sprinkler system or allows the fire doors to close automatically. One of the most common alloys with a low melting point is known as Wood's metal. It consists of 4 parts of bismuth, 2 parts of lead, 1 part of tin and 1 part of cadmium, and melts at a temperature much lower than the boiling point of water.

ACTIVITY SERIES OF THE METALS

In the discussion of metallic elements in the following pages, you will find certain metals referred to as active and others as inactive. The relative chemical activity of a metal depends upon the ease with which electrons are released from an atom of the metal. Since the release of electrons by the atom of a metallic element is just what occurs when the metal reacts with oxygen to form an oxide, the term "oxidation" has been broadened to cover all similar reactions. A metal, then, which loses electrons easily will combine readily with oxygen and also with other non-metallic elements. The accompanying table shows the relative activity of the metallic elements, starting with the most active and ending with the least active. Hydrogen is included in the table because, like the typical metals, it combines with oxygen and can undergo many other oxidation reactions by losing electrons. Any metal which is more active than hydrogen, that is, which stands above hydrogen in the table, can displace hydrogen in the form of gas from an acid solution. In previous experiments you used iron, aluminum, and zinc to generate hydrogen by reaction with acid, but copper and silver would not work, since they are less active than hydrogen, standing below it in the table.


Increasing chemical activity
Increasing tendency to lose electrons
Potassium
Sodium
Calcium
Magnesium
Aluminum
Manganese
Zinc
Iron
Cadmium
Nickel
Tin
Lead
HYDROGEN
Copper
Mercury
Silver
Platinum
Gold
Decreasing chemical activity
Tighter hold on electrons

ACTIVITY SERIES-
The best known metallic elements arranged in the order of their activity

THE ALKALI METALS

Group Ia of the Periodic Table is known as the alkali metals and consists of lithium, sodium, potassium, rubidium and cesium. The metals in this group all have the valence of one. Since these metals readily release an electron, they are chemically quite active and readily form a large number of compounds. The ions formed by the alkali metals are colorless and with but few exceptions their compounds are white and are readily soluble in water.

SODIUM

Most important among the alkali metals is sodium. In the metallic form, sodium is a silvery, soft, waxy substance which combines directly with water forming sodium hydroxide.

The chief source of sodium and its compounds is, of course, sodium chloride (common salt). All animal life requires sodium chloride and this salt is the principal raw material for practically all sodium compounds, among the most important of which are sodium hydroxide, sodium carbonate, sodium bicarbonate and sodium sulfate.

Sodium hydroxide is a strong base and because of its corrosive action, it is said to be "caustic." Sodium hydroxide is often called caustic soda. It is readily soluble in water, forming a strongly alkaline solution.

EXPERIMENT 214—Making Sodium Hydroxide.

Put 2 measures of Calcium Oxide (No. 20) and 2 measures of Sodium Carbonate (No. 4) in a test tube. Fill the tube ¼ full of water and boil the liquid for a minute or two, being careful that it does not boil over. Keep your face away from the tube while heating the liquid.

Allow the solution to cool and the solid matter to settle, then put a drop of the clear liquid on the end of your finger and rub it around, noticing the soapy feel. This is characteristic of strong bases like sodium hydroxide.

Sodium carbonate, sometimes called soda ash, is an important compound of sodium and is quite widely used in industrial processes, as well as in the home as washing soda.

EXPERIMENT 215—Making Sodium Carbonate from Sodium Bicarbonate.

Dissolve 1 measure of sodium bicarbonate (baking soda) in a test tube half full of water. Add 1 drop of Phenolphthalein Solution (No. 22) to the sodium bicarbonate solution and notice that a very faint pink color results.

Now carefully boil the solution for a few moments, and you will find that the pink color becomes a deep red color.

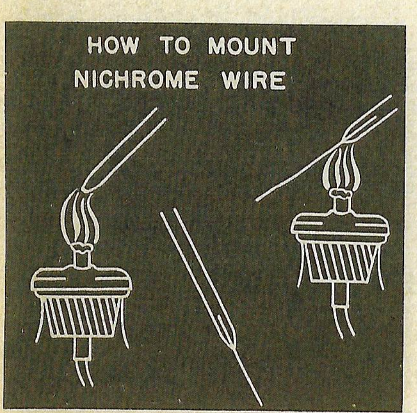
Sodium bicarbonate is a very mild alkali, but when heated it gives off carbon dioxide and changes into sodium carbonate which is a stronger alkali. Thus, when you heated the sodium bicarbonate solution you changed it into sodium carbonate solution.

Many metals and their salts possess the very interesting property of imparting characteristic colorations to a flame, such as that of an alcohol lamp or bunsen burner. This property makes it possible to identify certain metals in a chemical compound. To make these flame tests you will need a piece of Nichrome Wire (No. 58) fused into a glass tube handle.

EXPERIMENT 216—How to Mount your Nichrome Wire.

Heat the end of a piece of glass tubing, approximately six inches long, in the flame of the alcohol lamp, turning the tube constantly so that it is heated evenly. When it becomes soft, you will notice that the end of the tube is gradually closing. At this point remove the tubing from the flame and insert one end of the Nichrome Wire (No. 58) into the small opening for about one-half inch. Again heat the end of the tube until the wire is fused into the glass.

Allow the tube to cool and then fire polish the opposite end of the tube by holding it in the flame of the alcohol lamp until the rough edges are rounded.



EXPERIMENT 217—Flame Color Test for Sodium.

Moisten your nichrome wire and hold it in the flame of the alcohol lamp until the flame is no longer colored by the wire. Now again moisten the nichrome wire and touch it to some sodium chloride (common table salt), hold the wire in the alcohol lamp flame and notice the bright yellow color that is imparted to the flame. This is characteristic of sodium and sodium compounds. The sodium flame is very persistent and it may take quite a little time to burn all of the sodium off your nichrome wire. Dipping the wire in clear water occasionally and then holding it again in the flame will help to get rid of this color.

Try this test with other sodium compounds in your outfit, such as sodium bisulfate, sodium thiosulfate and sodium carbonate.

POTASSIUM

Potassium is very similar, both physically and chemically, to sodium. The pure metal is slightly more active than sodium metal and when a piece of potassium is dropped into water the reaction is so vigorous that the hydrogen liberated from the water catches fire and burns.

Potassium is not as plentiful as sodium although potassium compounds are very widely distributed throughout the earth. All growing plants need potassium and this element must, therefore, be in the soil in such a form that the plants can absorb it. Most mixed fertilizers contain potassium in the form of one of its soluble salts.

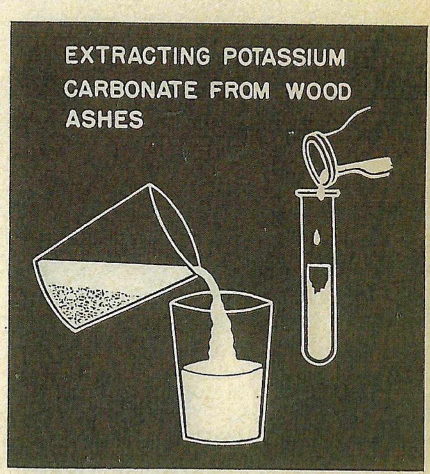
Potassium salts are obtained from many different sources, principally from deposits left by the drying up of lakes or inland seas. Seaweed contains a large per cent of potassium salts and potassium compounds are also obtained as a by-product from cement kilns. Formerly, the principal source of potassium was wood ashes.

EXPERIMENT 218—Extracting Potassium Carbonate from Wood Ash.

Put 2 or 3 spoonfuls of fresh wood ashes in a tumbler and fill it ½ full of water. Stir for a few minutes and allow the solid matter to settle to the bottom. Pour off some of the clear solution into another glass.

Test a small portion of this liquid in a test tube by adding a few drops of Phenolphthalein Solution (No. 22) and also by dipping into it a piece of red Litmus Paper (No. 18). You will find that it gives an alkaline reaction because of the potassium carbonate dissolved out of the wood ashes.

The product obtained by evaporating this solution of impure potassium carbonate is known as "pearl ash" and was formerly an important industrial product.



EXPERIMENT 219—Flame Color Test for Potassium.

Clean your flame test wire by heating it in the flame of the alcohol lamp until it no longer gives a color to the flame. Then dip it into the solution obtained from the wood ashes in the preceding experiment and heat it again in the flame. Look at the flame through a piece of blue glass and you will see the characteristic violet color given to a flame by potassium.

THE ALKALINE EARTH METALS

The elements of Group IIa in the Periodic Table are called alkaline earth metals because their oxides were called alkaline earths in the early days of chemistry. The metals comprising this group are beryllium, magnesium, calcium, strontium, barium, and radium. All of these metals have a valence of two.

As a group, the alkaline earth metals are second in activity only to the alkali metals. With the exception of beryllium, these metals are fairly soft.

The only metal of the alkaline earth group which is very useful in the metallic form is magnesium. This metallic element is very strong in proportion to its weight and therefore finds considerable commercial use where strength and extreme lightness are required.

CALCIUM

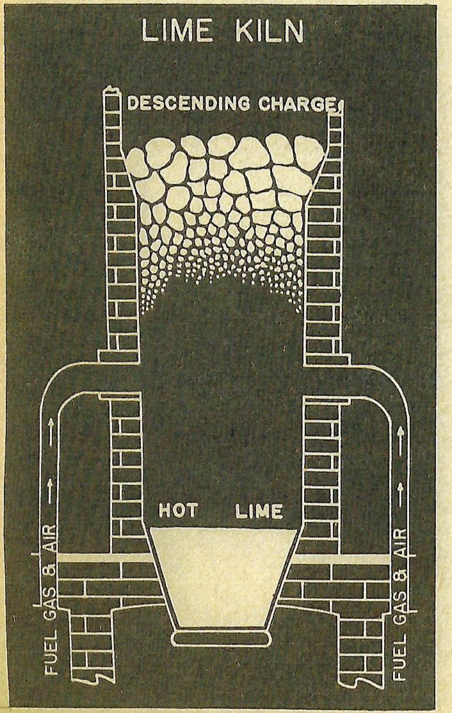
Calcium, which is a silvery white metal, decomposes water, forming calcium hydroxide and liberating hydrogen:

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2\uparrow$$

It is one of the most abundant elements and in combination with other elements forms nearly 3½% of the earth's crust. In combination with carbon and oxygen it forms calcium carbonate (limestone), with phosphorus and oxygen it occurs in phosphate rock, and with sulfur and oxygen it forms gypsum. Pearls are largely calcium carbonate, the same thing chemically as limestone.

A wide variety of calcium compounds are used in industry. Calcium oxide, or lime, is used for making mortar and plaster, and a complex compound of calcium with chlorine and oxygen is called bleaching powder, and is used in bleaching and as a disinfectant. Portland cement is made by heating a mixture of limestone and clay in kilns to an extremely high temperature and then pulverizing to a fine powder. The result is a mixture containing chiefly various calcium silicates and calcium aluminates.

Calcium oxide or lime is made by heating calcium carbonate.



EXPERIMENT 220—How to Make Lime.

If you can obtain some small pieces of limestone or marble, or even oyster shells, you can very easily make a small quantity of lime. Place 2 or 3 teaspoonfuls of the limestone or other material in the lid of a baking powder can or similar container. The limestone should be broken into rather small lumps.

The stone must now be heated to a very high temperature so it will be necessary to set the lid on a hot stove. Heat the limestone until it becames a bright red or nearly white color, that is, red hot, and then remove it from the fire and allow it to cool. You will find that it is different from the original limestone, in fact it is changed into calcium oxide or lime.

EXPERIMENT 221—Showing How Lime Expands When It is Slaked or Hydrated.

Fill a very small bottle or vial with finely ground lime. Pour a little water into the bottle and set the bottle aside to see what happens. As the lime slakes, it expands and occupies a greater volume. Since you have the lime confined in the bottle it will either have to break the bottle or come out of the mouth as it slakes. Be extremely careful when performing this experiment

EXPERIMENT 222—Flame Color Test for Calcium.

Moisten your Nichrome Wire (No. 58) and touch some Bleaching Powder (No. 38). This compound contains some calcium hypochlorite and when you hold the wire in the alcohol lamp flame, you will notice the yellowish red color, which is characteristic of calcium. The flame color test for calcium will not work with calcium oxide since the volatilization or boiling point of lime is 2850° C. This temperature cannot be attained in the alcohol lamp or bunsen burner flame.

MAGNESIUM

Magnesium is a silvery white metal which oxidizes to some extent in the air. Its compounds are widely distributed in the earth. Commercially, magnesium compounds are obtained from salt beds and natural brines. It is purified to the metallic form by electrochemical means. Powdered magnesium metal and magnesium ribbon burn brightly in the air and are used in fireworks. Magnesium is also used in flashbulbs for indoor photography. Magnesium is a constituent of duralumin and other light alloys which are important in airplane construction. Magnesium oxide, MgO, is used in lining high temperature furnaces and in making sorrel cement for stucco and floors. Milk of magnesia, used in medicine for neutralizing acidity, is a suspension of magnesium hydroxide in water. Epsom salts consist of magnesium sulfate with water of hydration, $MgSO_4.7H_2O$.

STRONTIUM

The compounds of strontium resemble those of calcium, both in physical properties and in chemical behavior. Most of our strontium is obtained from the mineral celestite, strontium sulfate. The carbonate of strontium, strontianite, is also found in nature. The pure metal is obtained by electrolysis of molten strontium chloride.

Strontium hydroxide is used in sugar refining, and strontium nitrate and strontium chloride are used in fireworks and in the manufacture of red flares and signals.

EXPERIMENT 223—Flame Color Test for Strontium.

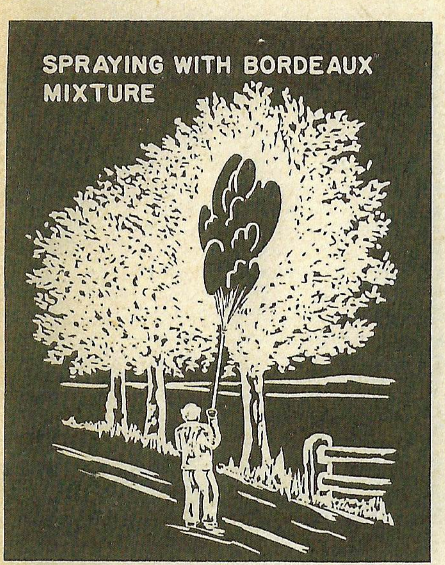
Moisten your nichrome wire and touch it to some Strontium Chloride (No. 10). Hold the wire in the alcohol lamp flame and notice the beautiful red coloration that is imparted to the flame. Strontium salts give this characteristic red color to flame.

COPPER, SILVER AND GOLD

The metals copper, silver and gold are placed in Group Ib of the Periodic Table. Each of these metals may have a valence of 1, 2, or 3. Copper usually exhibits a valence of 1 or 2. Silver usually has a valence of 1 and gold usually exhibits a valence of 1 or 3. They are generally encountered with a valence of one. All three of these metals are difficult to oxidize and they are, therefore, frequently found in the metallic state in nature and to a great degree resist corrosion. Copper is the most easily oxidized of the three metals. Silver is more difficult and gold cannot be oxidized by direct means. Due to their resistance to oxidation and their relative chemical inactivity, silver and gold are known as the "noble metals."

COPPER

Copper metal has been known and used since prehistoric times, principally because it often occurs in nature as the metallic element and its physical characteristics allowed it to be formed easily into weapons and utensils. Among the more important ores of copper are various sulfides, such as chalcopyrite and chalcocite, an oxide, suprite and a basic carbonate, malachite. The copper is separated from its ores by roasting followed by smelting, and the impure metal which is obtained is refined by the electrolytic process.



Copper is one of the few metals which has a distinguishing color. Owing to its property of conducting electricity, large quantities of this metal are used in the electrical industries. Copper is also used in equipment for chemical plants, in the building industries and in making electrotypes for printing. We have already mentioned some of the alloys of copper, such as brass and bronze, which are of great importance.

Since copper exhibits two valences, it forms two oxides, cuprous oxide, Cu_2O (valence 1) and cupric oxide, CuO (valence 2). Corresponding to these two oxides, there are two series of copper salts, the cuprous and cupric. Cupric salts are of great importance and cupric sulfate, $CuSO_4$, sometimes called bluestone, or blue vitriol, has use in many industrial and commercial applications. Dissolved in water and mixed with lime, cupric sulfate forms Bordeaux mixture, an important spray material for controlling fungus diseases of plants.

EXPERIMENT 224—Flame Color Test for Copper.

Moisten your nichrome wire and heat it until all traces of color to the flame have disappeared. Now again moisten the wire and touch it to Azurite (No. 39). Hold the wire in the flame again and you will note a very pretty bluish green color, which is characteristic of copper.

EXPERIMENT 225—Testing Brass for the Presence of Copper.

Hold a small piece of sheet brass (1" x ½") with a pair of pliers and heat it over the flame of the alcohol lamp until it becomes red hot. You will notice a beautiful green coloration in the flame, identical with that observed when heating a piece of copper wire. After heating the piece of brass for a few minutes, remove it from the flame and allow a few drops of cold water to fall upon the metal. A dark coating of black copper oxide will be formed, proving that brass contains copper.

SILVER

Like copper, silver occurs free in nature, often in the form of mixtures or alloys with other metals. Among the silver minerals, the sulfide, argentite, is the most important source of this metal. Silver is frequently associated with the sulfide minerals of copper and lead.

Silver is the best of all conductors of heat and electricity. It is not oxidized in the air, but as we have learned, it does combine with sulfur, forming silver sulfide. The principal commercial uses of silver are in silverware for the home, jewelry and for coinage. As we have learned, the silver coins of our country contain 90% silver. Sterling silver contains 92.5% silver. Since pure silver is soft, it is alloyed with small amounts of copper to harden it.

Among the more important compounds of silver are silver nitrate and the silver halides, including silver chloride, bromide and iodide, which are used in the photographic industry because of their sensitiveness to light.

EXPERIMENT 226—Oxidized Silver.

"Oxidized" silver is silver which has been given a rich, dark tone by the formation of a layer of silver sulfide on its surface. A wide variety of techniques is used for doing this. You can do it very simply by placing 1 measure of Sodium Thiosulfate (No. 16) on a bright silver coin and heating it gently in your alcohol lamp flame. Hold the coin by one edge with your test tube holder. First the solid melts and then steam comes from it (its water of hydration). Slowly the surface of the silver darkens as sulfur from the thiosulfate combines with it. Rinse the coin when it has become as dark as you want it.

EXPERIMENT 227—How to Clean Silverware with Sodium Thiosulfate.

It may seem strange that the same compound with which you darkened silver in the preceding experiment can also be used to clean it. In order to darken silver, you heated the sodium thiosulfate. Now using sodium thiosulfate in solution, you will see that it dissolves the silver sulfide. Prepare a solution of Sodium Thiosulfate (No. 16) in a test tube ¼ full of water. Moisten a clean soft cloth with this solution and rub a piece of silverware which is tarnished with black silver sulfide. You will find that the silver sulfide comes off very easily.

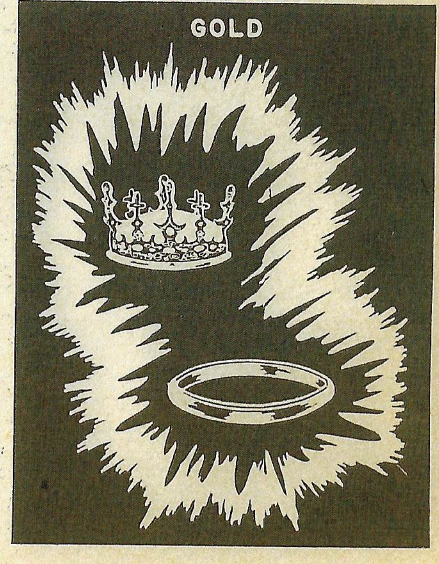
EXPERIMENT 228—The Electrolytic Cleaning of Silverware.

Obtain a piece of silver which is stained dark with silver sulfide. (If necessary expose the silver to some hydrogen sulfide gas.) Place the silver in an aluminum pan, cover with water and add common salt at the rate of one teaspoonful to each quart of water. Heat the pan on the stove until the water boils and the tarnish disappears from the silver. Only a minute or two of boiling should be necessary.

The aluminum is corroded while the silver is cleaned, so it is best not to leave the silver in the pan longer than necessary. If an old leaky aluminum pan is available, this may be used by setting it in a large enameled or agateware pan.

GOLD

From the beginnings of history, gold has been held as the symbol of wealth and as the king of metals. Its beautiful color and luster, its freedom from corrosion, the ease with which it is worked and shaped and finally, its rarity have caused it to be most highly prized. In more modern times gold has become the security for the monetary systems of the greater part of the world.



As might be expected, from its chemical inactivity, gold is generally found in the free state in nature, although often somewhat contaminated with other metals, especially silver. It occurs as veins in quartz rock and as this rock is disintegrated and washed down by streams, the gold accumulates in the stream bed as chunks called nuggets, and as flaky particles. Nuggets of gold weighing 100 lbs. or more have been discovered, but much more numerous are the very small fragments, ranging down to the size of dust specks.

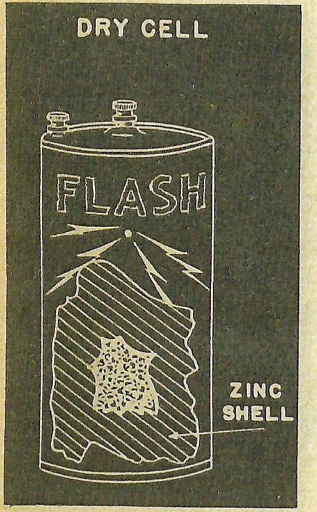
Gold is the most malleable of all metals. With care, it can be hammered down to a thickness of three or four millionths of an inch. It is not attacked by any single acid, but a mixture of nitric and hydrochloric acids, known as aqua regia, dissolves gold, converting it into the chloride, known as auric chloride, $AuCl_3$.

The principal uses for gold are coinage, jewelry, plating for ornamental work and in dental work. As pure gold is quite soft, it is usually alloyed with silver and copper to make it harder. Pure gold is 24 carats fine. Thus 18 carat gold would be 18/24 pure or 75% gold.

ZINC, CADMIUM AND MERCURY

Zinc, cadmium, and mercury constitute sub-group b of Group II of the periodic system, being thus cousins of the alkaline earth metals, to which they bear about the same relation as copper, silver, and gold bear to the alkali metals. These three elements all show a valence of two, but mercury sometimes has in addition a valence of one. Since zinc and cadmium resemble each other much more than mercury, we may discuss them together.

ZINC AND CADMIUM



The most common zinc minerals are the sulfide, ZnS, called zinc blende or sphalerite; the carbonate, $ZnCO_3$, called smithsonite; and a complex combination of the oxides of zinc, iron and manganese, known as franklinite. Cadmium is found in many zinc ores in small quantities.

In obtaining metallic zinc, the ores are roasted to convert them into zinc oxide. The oxide is heated with powdered coal, thus being reduced to metallic zinc, which comes off in the form of a vapor and is condensed in a receiver.

Zinc is a silvery white metal, but soon becomes bluish gray, due to the formation of a film of basic carbonate, formed by the action of oxygen and carbon dioxide in the air. At high temperatures, zinc burns in the air to form zinc oxide. It displaces hydrogen easily from acids with the formation of the corresponding salts of zinc.

Cadmium is very similar in physical and chemical properties to zinc.

The principal use of zinc is as a coating on iron and steel (galvanizing). This coating protects the iron from corrosion. Zinc is also used quite widely in various alloys, especially with copper. Metal for diecasting consists principally of zinc. Zinc is used in dry cell batteries, the shell or container of each cell often being made of zinc and serving as the negative electrode. Large amounts of zinc oxide are used as pigment for white paint.

Cadmium is extensively used as a protective coating on iron and steel. It is equal in protective power to zinc and it makes a better looking finish. Cadmium is also alloyed with other metals to make bearings and it is used in making the low melting alloys used in automatic sprinkler heads.

EXPERIMENT 229—Dissolving Zinc in Vinegar.

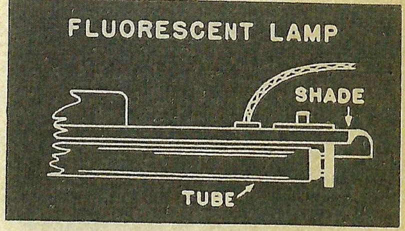
Fill a test tube ¼ full of ordinary vinegar and drop in a piece of zinc metal about the size of the end of your measure. Warm the vinegar gently to speed up the reaction. The bubbles which rise from the zinc are hydrogen gas. Set the test tube aside for an hour or two to give the acetic acid of the vinegar plenty of time to dissolve the zinc.

Meanwhile dissolve in another test tube ¼ full of water 1 measure of Sodium Carbonate (No. 4), and in a third test tube also ¼ full of water, dissolve 1 measure of Sodium Ferrocyanide (No. 6). After sufficient zinc has reacted with the vinegar, pour a few drops of the zinc solution into each of the other two test tubes. In the sodium carbonate solution you will find that the solution containing zinc acetate gives a white precipitate of basic zinc carbonate. In the sodium ferrocyanide solution a precipitate of white zinc ferrocyanide is formed. The fact that these precipitates are formed shows that zinc went into solution in the vinegar. This experiment demonstrates why galvanized utensils are not suitable for cooking acid foods.

MERCURY

Mercury is the only metallic element which is liquid at ordinary temperatures, although gallium is a liquid just above the usual room temperature. Mercury is heavier than lead, having a density of 13.6 grams per cubic centimeter. It dissolves many metals with the formation of alloys called amalgams. Some of these are solid and some are liquid at ordinary temperatures.

Mercury is sometimes found in the free state in nature often amalgamated with silver or gold. The most important source of mercury, however, is the mineral cinnabar, which is mercuric sulfide. This ore is roasted in the presence of air so as to burn out the sulfur and the metallic mercury distills off in the form of vapor and is condensed in a receiver.



The unique properties of metallic mercury make it useful for a great variety of purposes. Its application in thermometers and barometers is familiar. Amalgams of silver, gold and tin are used in dental fillings. When electricity is passed through mercury vapor at low pressure, light is emitted, which is very rich in ultra violet radiation. Thus mercury vapor lamps in a quartz tube, which is transparent to ultra violet light, are an important source of ultra violet radiation in chemical and therapeutic operations. Certain compounds are capable of absorbing the ultra violet radiations and re-emitting this light energy in visible form. Thus, if we coat the inside of a tube with such substances, called phosphors, fill the tube with mercury vapor and pass a current through it, we have the so-called fluorescent lamp, which is widely used in modern lighting.

Soluble mercury compounds are highly poisonous, for example, mercuric chloride, often called bichloride of mercury. Mercurochrome is a complex organic mercury compound which is used as an antiseptic on cuts and wounds.

EXPERIMENT 230—Mercury from Mercurochrome.

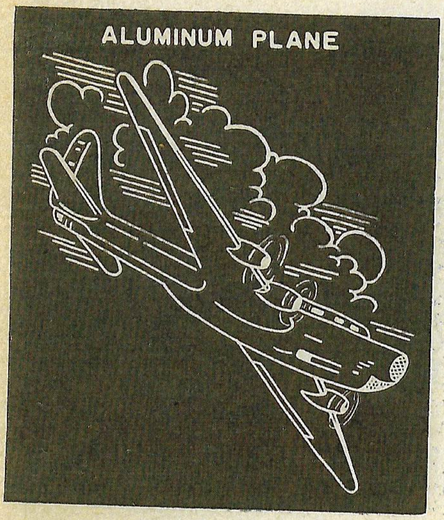
Fill a test tube ⅛ full of Mercurochrome solution and add an equal amount of water. Drop in 2 measures of Sodium Bisulfate (No. 7) and heat to boiling. Then add ½ measure of Bleaching Powder (No. 38) and agitate the test tube to mix in the bleaching powder. Continue to heat gently, being very careful not to let the slushy mixture spurt out. Soon the color begins to fade and after several minutes, it should be almost entirely gone. If necessary, add one or more additional small portions of bleaching powder, but avoid putting in a large excess. Mercurochrome is a complex organic mercury compound and this treatment decomposes it and brings the mercury into the form of ions. When most of the colored material has been decomposed, pour the solution into a tumbler and immerse in it a clean, bright penny. The penny may be cleaned with steel wool or by boiling it in vinegar. After a few moments, the penny becomes silvery due to the mercury which deposits upon it, forming an amalgam with the copper. Rinse the penny and rub it with a soft cloth to make it shiny.

ALUMINUM

Aluminum is the most important metal in Group IIIa of the Periodic Table. The other metals in this group and also the metals in Group IIIb are all rarer elements which are of little importance commercially. Aluminum is trivalent, meaning that it has a valence of three.

Aluminum is the third most plentiful of the elements in the crust of the earth, the first being oxygen and the second, silicon. Aluminum is contained in many of our common mineral substances, including clay, feldspar, mica, garnet, corundum and bauxite. This latter, bauxite, is the chief commercial ore of aluminum and is a form of aluminum oxide.

In spite of the abundance of aluminum, the great chemical activity of the free metal makes its extraction in the pure state quite difficult. The first practical method of obtaining aluminum for its ores was discovered in 1886 by Charles M. Hall, a student at Oberlin College. The purified oxide of aluminum is melted in cryolite (a sodium aluminum fluoride) and an electric current passed through the solution. The melting point of aluminum oxide is too high to permit melting it directly for electrolysis, but it readily dissolves in melted cryolite and the aluminum ions in this solution are set free as aluminum metal in the cathode of the electric cell.



The most important use for aluminum is in the metallic state and as you know, many cooking utensils and articles about the home are made of aluminum. Industrially, aluminum is used in the construction of airplanes, trains, buses, truck bodies, or wherever light weight, strength and permanence are of importance.

Among the important compounds of aluminum is aluminum oxide from which the metal is made and which is also used in the manufacture of abrasives, either in the natural form known as corundum, or in an artificial form known as alundum. Aluminum oxide is also used as a refractory in lining high temperature crucibles and in making firebrick for industrial furnaces. Rubies and sapphires are composed of aluminum oxide, colored by traces of the oxides of other metals. Aluminum sulfate is used quite widely in clarifying water. Aluminum also forms a class of compounds known as alums, which are really double sulfates. Potassium alum, for example, is $K_2SO_4.Al_2(SO_4)_3.24H_2O$. Sodium alum is $Na_2SO_4.Al_2(SO_4)_3.24H_2O$. And ammonium alum is $(NH_4)_2SO_4.Al_2(SO_4)_3.24H_2O$.

EXPERIMENT 231—Dyeing Aluminum.

Recently it has been found possible to color aluminum with dyes by first forming a suitable coat of aluminum oxide on the surface.

Dissolve ½ measure of Sodium Carbonate (No. 4) in a test tube ¼ full of water and add 10 or 12 drops of ordinary hydrogen peroxide. Dip into this solution a small piece of aluminum (cut from an old utensil) and boil several minutes. This will produce a thick coating of aluminum oxide on the metal.

In another test tube place 5 or 6 drops of blue ink and fill the tube ¼ full of water. Now remove the aluminum strip from the oxidizing bath, wash it in running water and immerse it in the dye solution. After 5 minutes take it out, wash again in running water and let it dry.

EXPERIMENT 232—Dyeing Aluminum Red.

Repeat the preceding experiment, but instead of the blue ink, use ½ test tube full of red ink. You will find that the strip of aluminum metal takes up red coloring matter from the ink.

EXPERIMENT 233—Dyeing Aluminum Green.

By using green ink you will find that the aluminum oxide takes up green dye where the oxide is sufficiently thick.

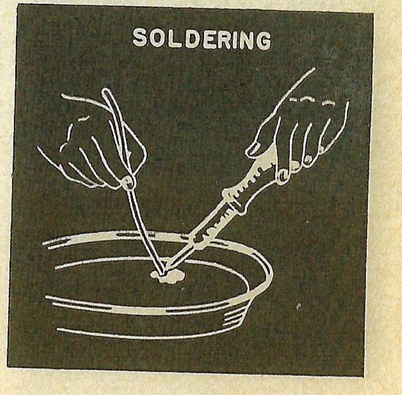
TIN, LEAD AND OTHER METALS OF GROUP IV

Tin and lead, together with germanium constitute Group IVb of the Periodic Table, while titanium, zirconium, hafnium and thorium constitute Group IVa. The important metals in Group IV are, of course, tin and lead. Both of these metals may possess a valence of either 2 or 4. Tin and lead are typically metallic from the standpoint of physical properties.

TIN

Tin is a relatively scarce metal and unfortunately, the United States has no important deposits of tin ores. The only important ore of tin is cassiterite, or tinstone, which is stannic oxide, SnO_2. The most important sources of this ore are Malaya, the Dutch East Indies and Bolivia. In producing metallic tin, the ore is roasted and then smelted or reduced with carbon in the form of coal or coke,

$$SnO_2 + 2C \rightarrow Sn + 2CO$$



Tin is a relatively inactive metal and hence is rather easily obtained from its oxide. Conversely, its oxide is not very readily formed, and so tin is fairly resistant to corrosion.

The greatest use of metallic tin is in coating steel to protect it from corrosion. Our so-called "tin" cans contain only 1% or less of tin deposited in an extremely thin film over the surface of the steel. This coating was formerly done by passing the steel through a bath of melted tin. In order to save tin, however, most of the coating is now done by electroplating. Tin is also of importance in many alloys. Solder contains various proportions of tin and lead. Type metal consists of lead, antimony and tin. Babbitt metal used in bearings to reduce friction contains copper, antimony and tin.

LEAD

Lead is a soft heavy metal having a specific gravity of 11.3. The principal source of lead is lead sulfide, PbS, the mineral galena. The concentrated lead ore is roasted to burn out part of the sulfur, leaving a mixture of lead oxide and lead sulfide. This is mixed with coke and limestone and charged into a blast furnace. Lead sulfide and lead oxide react, setting free the lead. The coke also acts as a reducing agent. The purpose of the limestone is to combine with any silica which may be present in the ore.

Due to its resistance to corrosion and also to the plentiful supply and low price, lead has a large number of uses. The acid type of storage battery uses lead plates. Telephone cables are sheathed in lead to protect them. Much equipment in chemical plants is constructed with lead or lined with it since it is resistant to chemical attack. Lead shot and bullets are familiar. Lead pipe is used extensively for drains. Lead is used in a number of alloys, as previously mentioned.

Some of the compounds of lead are of great importance industrially. Lead monoxide, PbO, in which the valence of lead is 2, is commonly called litharge. Litharge is used in paints as a drying agent and it also forms a waterproof cement when mixed with glycerin. Lead dioxide, PbO_2, in which lead has a valence of 4, is the active substance on the positive plate of a charged storage battery.

White lead is one of our most valuable white pigments for paint. It is a basic carbonate, that is, a combination of lead hydroxide and lead carbonate. Red lead, Pb_3O_4, is commonly used in paints for structural steel. Lead arsenate is an important insecticide, and tetraethyl lead is the anti-knock fluid added to ordinary gasoline to make ethyl gasoline.

ARSENIC, ANTIMONY AND BISMUTH

Arsenic, antimony and bismuth are classified in Group V, sub-division-b of the Periodic Table, while vanadium, columbium and tatalum are classified in Group Va. The most important metals in these groups are arsenic, antimony and bismuth. These metals each possess a valence of either 3 or 5 and in addition, antimony may in rare cases be 4 valent. These three elements, although they look like metals, possess many of the properties of non-metals. The oxide of arsenic, for example, when dissolved in water, forms an acid instead of a base. The oxides of antimony and bismuth, however, when dissolved in water, form weak bases. The property of arsenic mentioned above accounts for the fact that its oxide can combine with bases, forming arsenates.

ARSENIC

Free arsenic has few uses, the chief being as an alloying constituent with lead in making shot where it tends to harden the lead. By far the greatest use for arsenic compounds is in the preparation of poisons for rodents and insects. The largest amounts of arsenic probably go into the manufacture of lead arsenate and for calcium arsenate for use in insect sprays.

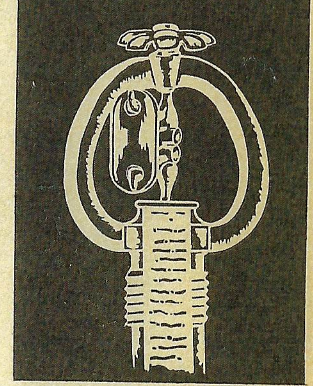
ANTIMONY

The greater metallic character of antimony is reflected in its use in alloys to which it imparts the very valuable property of expanding slightly upon solidifying. This is especially important in the casting of type for printing since it causes the solidified metal to fill the mold completely, giving a sharp impression. Like arsenic, antimony has the effect of hardening lead and also of increasing its resistance to corrosion. Hence, lead storage battery plates and lead sheaths for telephone cables contain some antimony. The sulfides of antimony, when prepared by precipitation, are orange in color and are used as a pigment in coloring rubber goods.

BISMUTH

The most important and interesting use of bismuth is in low melting point alloys. The melting point of solid bismuth is 271° C., but when we dissolve in the bismuth 25% of lead and 12½% each of tin and cadmium, the melting point of the alloy is lowered to 70° C., much below the boiling point of water and far below the melting point of any of the added metals. This alloy is used in the heads for automatic sprinkler systems, arranged so that when the alloy is melted by excessive heat, a valve is released and the water is turned on.

SPRINKLER HEAD



CHROMIUM, TUNGSTEN, URANIUM AND MANGANESE

In this section we shall discuss several less abundant, but nevertheless, important metals. Among them are representatives of several groups of the Periodic Table and the reactions of most of them are rather complicated since they may exist in several valence states.

CHROMIUM

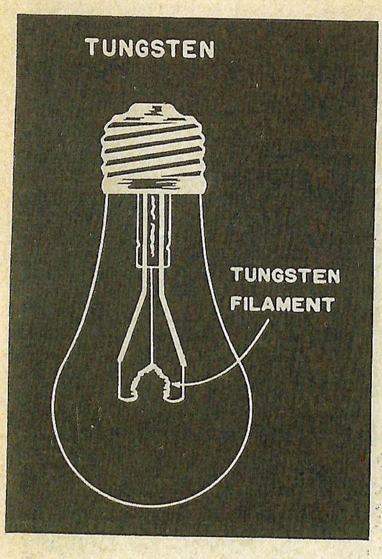
Chromium is classified in the Sixth Group, sub-division a, of the Periodic Table and its valence may be either 2, 3, or 6.

The most important use for chromium is in alloys with steel to prevent corrosion. Typical stainless steel contains about 18% chromium along with 8% nickel. Chromium is also electroplated on various metal parts. It produces a bright surface of great hardness and strength, which is highly resistant to corrosion.

The most important compounds of chromium are the chromates and dichromates. These compounds are the basis of certain yellow and orange pigments used in paints and they are also used in tanning leather, as chrome-tanned hides make the most waterproof leather.

TUNGSTEN

This element, together with molybdenum, is classified in the same group in the Periodic Table as chromium. Both of these elements are of great importance in alloy steels. Tungsten has the highest melting point of any



metal — 3370° C. As obtained by reduction of its oxide, tungsten is a fine powder. In order to form this metal into wire, it is heated under great pressure which causes it to sinter together sufficiently so that it can be worked into thinner and thinner rods and finally drawn into wire. This wire is used in incandescent lamp filaments and in radio and x-ray tubes.

URANIUM

Uranium is classified in the Periodic Table among the Rare Earths. It has the highest atomic number of any element found in nature.

The principal use for uranium was formerly in uranium glass, to which it imparts a yellowish color. Due to the remarkable reactions which its nucleus undergoes, uranium is now in large demand as a source of atomic energy.

MANGANESE

Manganese is classified in Group VII, sub-division a, of the Periodic Table. It may exist in five valence states, namely, 2, 3, 4, 6, or 7 valent. This results in the formation of five well defined sets of compounds arising from the five oxides of manganese. The lower oxides of manganese, MnO and Mn_2O_3, are basic, while the higher oxides, MnO_3 and Mn_2O_7, when dissolved in water are acid. The intermediate oxide, MnO_2, called manganese dioxide, is practically neutral, but shows very weakly acidic properties.

The most important compound of manganese is manganese dioxide, in which state it occurs in nature as the mineral, pyrolusite. Manganese is used as an alloy with steel, as it has the property of making steel very hard and tough.

An important use of manganese dioxide is in the production of ordinary dry cells, where it serves as a depolarizer to prevent, by its oxidizing action, the accumulation of hydrogen on the carbon pole of the cell.

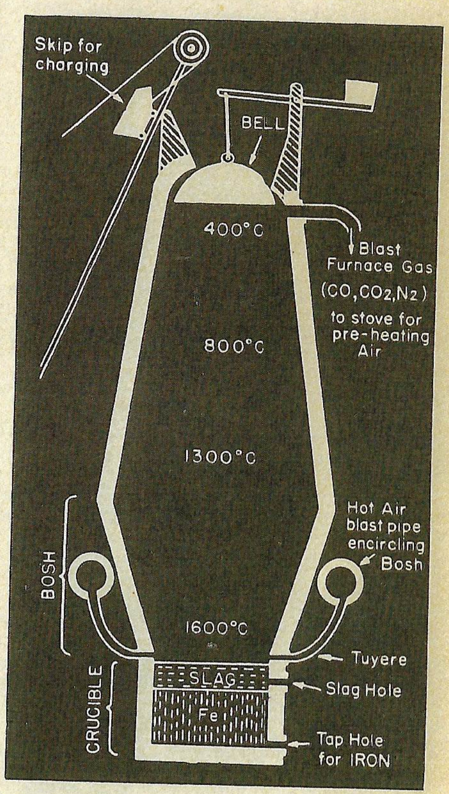
IRON, COBALT, NICKEL AND THE PLATINUM METALS

The metallic elements considered in this section are known as Transition Elements and appear under this heading in the Periodic Table. Iron, the most important metal in this group, commonly exhibits a valence of 2 or 3, but may in rare cases have a valence of 4 or 6. Cobalt likewise is commonly either 2 or 3 valent but may exhibit a valence of 4. Nickel is commonly 2 valent but in certain cases may show a valence of 3 or 4. Platinum exhibits a valence of 2 or 4 and may in rare cases have a valence of 3 or 6. Iron, cobalt and nickel resemble each other very closely.

IRON

Iron is the fourth most abundant element in the crust of the earth, following oxygen, silicon, and aluminum. It is thought that the core of the earth may be composed almost entirely of iron and nickel.

The most important ores of iron are all oxides and the most abundant of these is the mineral hematite, Fe_2O_3. Iron is obtained from its ores by a process of reduction carried out in a blast furnace. The reducing agent is carbon or rather the carbon monoxide obtained from it. The blast furnace is a great steel tower, about 100 feet high and 20 feet across, lined with firebrick. Charges of iron ore, coke and limestone are added through the top, while preheated air is blown in at the bottom. The reaction is one of reduction in which the iron oxide is reduced to metallic iron and the carbon, represented by the coke, is oxidized by the oxygen contained in the iron ore plus the oxygen in the air. Limestone is used to react with the silica and alumina which are impurities in the ore and forms a slag consisting of calcium silicate and calcium aluminate. The molten iron metal is drawn off at the bottom of the furnace.



Cast iron results when the impure product of the blast furnace is melted with scrap iron. This cast iron is relatively brittle but because of its cheapness, it is used for heavy machine bases, automobile cylinder blocks, steam and hot water radiators and many similar objects. Cast iron contains about 2% of carbon and part of this carbon is in the form of graphite interspersed with the grains of iron. The rest of the carbon is combined with the iron in the form of iron carbide.

EXPERIMENT 234—To Show the Presence of Free Carbon in Cast Iron.

Place ½ measure of powdered iron in a test tube. Add 6 measures of Sodium Bisulfate (No. 7) and fill the test tube ⅓ full of water. Warm gently until the sodium bisulfate is completely dissolved.

You will notice that the iron is being dissolved by the sodium bisulfate solution and that hydrogen gas is driven off. Mixed with this hydrogen is a small amount of hydrocarbon gases having an unpleasant odor. These gases are formed from the small amount of combined carbon in the cast iron. The free carbon or graphite, however. is not affected by the sodium bisulfate and remains as a black powder after all the iron has been dissolved.

STEEL

The term steel covers a number of different types of iron alloys and it is difficult to give an accurate definition. One of the reasons why iron is so useful is that its properties may be modified through a wide range by the use of small quantities of alloying elements. Chief among the elements which affect the characteristics of steel is carbon. When the carbon content is low (less then ½%), the steel will be relatively soft. Higher carbon content makes it possible to harden steel through heat treatment.

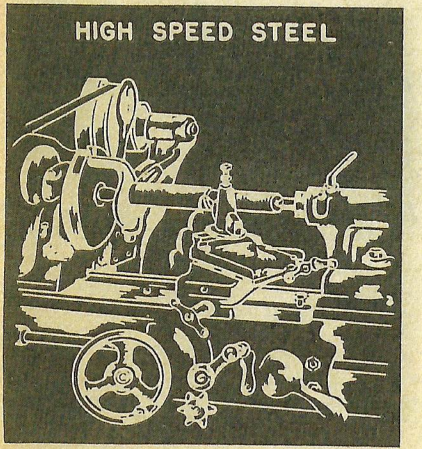
EXPERIMENT 235—Heat Treatment of Steel.

Heat a discarded safety razor blade to a bright red heat in the flame of an alcohol lamp or of a gas stove and gradually withdraw it from the flame so that it cools slowly. Then try bending it. You will find that it has lost most of its springiness. Suitable tests would show also that it is no longer very hard. You have "drawn the temper" of the steel.

Now heat the blade again, but this time cool it very suddenly by quenching in a vessel of cold water. After this treatment, you will find that the blade is extremely brittle and can be bent scarcely at all without breaking. It is, nevertheless, very hard. In making razor blades and similar articles, they are first hardened in this way by quenching and then heated to a carefully controlled temperature and allowed to cool slowly so as to soften them just enough to remove most of the brittleness. This also makes them less hard, of course, but some of the desirable hardness must be sacrificed for the sake of reducing the brittleness. As you did the experiment, a considerable amount of oxide scale was probably formed. To avoid this in commercial operations, the heating is frequently done in the presence of non-oxidizing gases.

Now try hardening, by the same process, a thin iron nail and see whether you can harden it enough so that it breaks when you bend it. You will probably not succeed, since nails are made of low-carbon steel, and unless at least 0.5% carbon is present, the steel can not be hardened by heat treatment.

Manganese will make steel tougher and harder. Nickel and chromium also give steel greater toughness and hardness. Relatively large amounts of nickel or chromium or both (up to about 20%) will produce stainless steel, which is not subject to oxidation or corrosion. A fraction of a per cent of vanadium increases the tensile strength of steel and is often used in spring steel. Tungsten makes steel very hard, even when it is red hot, so tungsten is used in so-called high-speed steels for making tools for machining metals. Usually chromium, vanadium and cobalt are used in connection with tungsten. Other elements sometimes used in alloy steels are copper, silver, lead, titanium, columbium, cerium, uranium, aluminum, tantalum, and no one can foretell what other elements may be found to give beneficial properties to steel, the backbone of our civilization.



You are familiar with several of the compounds of iron which have been used in previous experiments. The compounds of this metal, however, do not approach in importance the metal itself.

COBALT AND NICKEL

These metallic elements occur in the crust of the earth only to the extent of a small fraction of 1%, however, as mentioned, large quantities of nickel may be present in the center of the earth. As we have learned, both nickel and cobalt are of importance in steel where they impart special characteristics to the metal. Cobalt compounds are used to color glass blue and to impart a beautiful color to pottery glazes.

Nickel is used extensively in electroplating as metallic nickel does not oxidize or tarnish under ordinary conditions and it forms a hard bright durable finish on other metals. The principal use for nickel compounds is in plating baths where the nickel ions move toward the negative electrode and deposit on the work to be plated.

PLATINUM

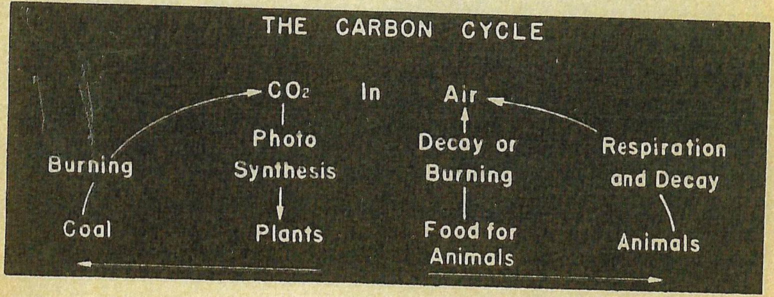
Platinum is the principal metal in a group of six metallic elements, all of which have quite similar physical and chemical properties. These elements are ruthenium, rhodium, palladium, osmium, iridium and platinum.

Platinum is used in jewelry, in dentistry and in chemical industries and laboratories. Platinum crucibles and evaporating dishes are indespensable because of their resistance to chemical attack. In finely divided form, platinum is a valuable catalyst in the contact sulfuric acid process and in the oxidation of ammonia to nitric acid.

Life On Our Earth

We live in a chemical world and we ourselves, as well as plants, animals and every living thing, depend on chemistry. We have learned that as we breathe, the oxygen of the air unites with carbon in our bodies forming carbon dioxide, which we exhale in our breath. Plants, on the other hand, absorb carbon dioxide from the air and give off oxygen, thus, both plant and animal life exist by chemical reactions.

The fact that all animal life, including man, adds to the supply of carbon dioxide in the air and that all combustion of coal, oil, gasoline and other fuels also adds to the supply of carbon dioxide, might lead us to believe that eventually the oxygen in our atmosphere would be all used up and that animal life could no longer exist. Counterbalancing this enormous use of oxygen and production of carbon dioxide is the absorption of carbon dioxide by plants which retain carbon and release oxygen to the air. Thus, animal and plant life are mutually dependent upon each other to maintain a proper balance in our chemical world.

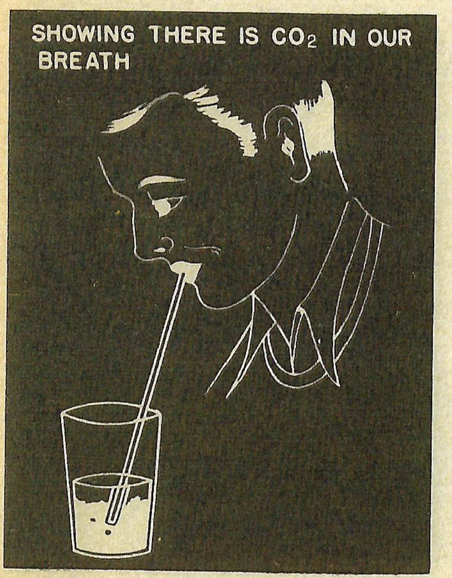


The leaves of green plants have the power of combining carbon dioxide and water to form sugars, starches, and cellulose. This reaction cannot take place unless energy in the form of light is supplied to the leaf, and hence the process by which plants use carbon dioxide and water is called "photosynthesis." Without this reaction animal life as we know it would be impossible. It was by means of photosynthesis that the energy of the sun's rays millions of years ago was stored in the form of coal. The carbon in coal was originally obtained from the carbon dioxide of the air by the energy of sunlight shining on green leaves.

Thus plants convert the carbon dioxide of the air into substances which can be used for food or fuel, and then the food is oxidized in the bodies of animals and the fuel is burned in a furnace, converting the carbon back again to carbon dioxide. The diagram shows the more important stages in this cycle chain of interconversions.

We have already learned in a previous experiment that leaves give off oxygen in the sunlight. Now let us demonstrate that carbon dioxide is present in the breath we exhale.

EXPERIMENT 236—Showing That There Is Carbon Dioxide in Our Breath.


SHOWING THERE IS CO_2 IN OUR BREATH

Fill an ordinary tumbler or drinking glass 1/4 full of water and add 1 or 2 drops of Phenolphthalein Solution (No. 22). In a clean test tube dissolve 1 measure of Sodium Carbonate (No. 4) in half a test tube of water. After the sodium carbonate is all dissolved, add 2 or 3 drops of the solution in the test tube to the glass containing the phenolphthalein and water. Add just enough of the sodium carbonate solution to turn the solution in the glass a pink color. It will take only a few drops.

Now, using your glass tube, blow into this pink solution in the glass. Inhale through your nose and exhale through the glass tube so that your breath bubbles through the pink solution. After you have done this for a few minutes, you will notice that the pink solution is gradually becoming lighter and after a while it will be entirely colorless. What has happened?

Since sodium carbonate is the salt of a very weak acid, carbonic acid, its solution is alkaline and that is why it turned the phenolphthalein solution pink. We also know that carbon dioxide is an oxide of a non-metal (carbon) and that when this non-metallic oxide is dissolved in water, it forms an acid (carbonic acid). Thus, as we blow our breath through the sodium carbonate solution, the carbon dioxide gradually dissolves in the water, forming carbonic acid. This acid converts the alkaline sodium carbonate into sodium bicarbonate, which contains a sufficient excess of the weak acid to make its solution nearly neutral. Thus, the phenolphthalein becomes colorless.

EXPERIMENT 237—Another Way of Proving the Presence of Carbon Dioxide in Our Breath.

Prepare some lime water by adding 2 measures of Calcium Oxide (No. 20) to a glass half full of water and stirring for a minute or two. Let the undissolved solid particles settle to the bottom and pour the clear solution into another glass.

Using a piece of glass tubing, blow through the lime water. You will find that as your breath bubbles through the liquid, it causes it to turn white and turbid, thus showing the presence of carbon dioxide.

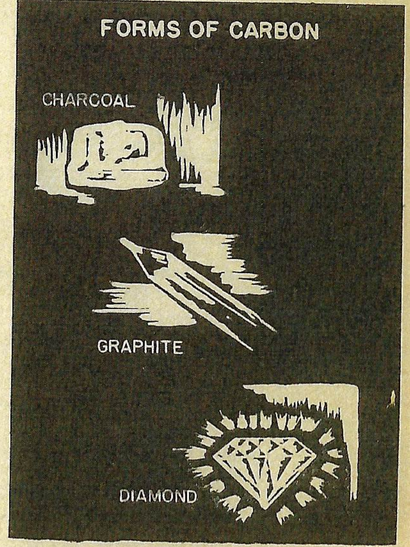
The white precipitate is calcium carbonate, which is soluble in most acids. Stir in a little vinegar and watch the turbidity disappear.

CARBON

Carbon is the first element in Group IVa of the Periodic Table. It may exist in the valence state of 2 or 4, but almost always exhibits a valence of 4. Carbon is classified as a non-metallic element because its oxide unites with water forming an acid.

Carbon, although it makes up less than a fifth of 1% of the earth's crust, is one of the most important elements because it is a constituent of so many valuable compounds and because it enters into numerous important reactions. It occurs in the free state as diamonds and as graphite. Many carbonates occur as rock-forming minerals, especially limestone and dolomite. Coal and petroleum are essentially mixtures of carbon compounds and all organic matter, plants and animals, consists in large part of carbon compounds. Furthermore, we find carbon in the form of carbon dioxide distributed through the atmosphere and dissolved in the water of the oceans.

Free carbon may exist in three allotropic forms. You will recall that sulfur and phosphorus exhibit similar properties. The allotropic forms of carbon are the diamond and graphite (both of which are crystalline in nature) and amorphous carbon, which may occur as coke, charcoal or lampblack.



EXPERIMENT 238—Forms of Carbon.

Examine a piece of graphite in a soft lead pencil. Notice its softness and smooth, greasy feel. It can be scratched even with your fingernail and rubs off onto paper, leaving a black smudge.

Now compare a piece of charcoal or coke with the graphite. It is harder and feels rough and harsh, although, it, too, will leave a mark on paper. Notice also that the charcoal is much less dense because of its many minute pores filled with air.

If you have a diamond ring, you can compare the most beautiful form of carbon with the other two. The transparent sparkle of the diamond is quite different from the dirty blackness of graphite and charcoal. Draw a corner of the diamond over a piece of glass. Notice how readily the glass is scratched. Graphite is one of the softest minerals. The diamond is the hardest of all, yet both are forms of pure carbon.

COMPOUNDS OF CARBON

Carbon forms many more compounds than any other element. Hundreds of thousands of compounds of carbon are already known and new ones are being discovered continually. It is no wonder then that the chemistry of the compounds of carbon has been set aside as a separate and very important branch of chemistry—organic chemistry. This tremendous variety of compounds formed from carbon is possible because carbon atoms, unlike those of most elements, are able to unite with one another, building up long chains which may then join into rings or have intricate branches attached to them, with other elements joined on at various places. These complicated carbon compounds studied by the organic chemist are of the very greatest importance. Some of them are fundamental to all living things. Others are dyes, drugs and explosives. Foods, textiles, wood, paper, rubber, plastics, coal and oil are all composed of organic compounds. We will first consider some of the simpler compounds of carbon, which are studied in inorganic chemistry.

CARBON MONOXIDE

When carbon burns in a limited amount of air, each atom of carbon unites with one of oxygen to form carbon monoxide:

$$\underset{\text{(carbon)}}{2\,C} \quad + \quad \underset{\text{(oxygen)}}{O_2} \quad \rightarrow \quad \underset{\text{(carbon monoxide)}}{2\,CO}$$

This compound is a gas, and it is extremely poisonous when inhaled. Since it is present in small amounts in the exhaust gases from automobile engines, it is well to be careful about breathing such gases, especially in closed garages. In particular, never run an automobile engine without opening the garage doors. Many people have lost their lives by disregarding this simple rule.

CARBON DIOXIDE

When carbon burns with an adequate supply of oxygen, each atom combines with two atoms of oxygen, and carbon dioxide results:

$$C + O_2 \rightarrow CO_2$$

It is also formed when carbon monoxide burns, taking up another atom of oxygen. A more convenient way of making it, however, is to treat a carbonate with an acid.

EXPERIMENT 239—Making Carbon Dioxide.

To a test tube ⅔ full of water, add ½ measure of Calcium Oxide (No. 20) and shake it vigorously for a few minutes. Let the undissolved solid settle to the bottom and pour off the clear liquid into a third test tube. This is lime water, a solution of calcium hydroxide.

Place 3 measures of baking soda in a test tube. Add ¼ test tube of vinegar and quickly fit the gas delivery tube to the mouth of the test tube, insert the long end of the delivery tube into the lime water, and let the carbon dioxide gas which is set free from the baking soda pass through the tube into the lime water. When the evolution of gas slackens, remove the stopper for an instant and drop in another measure of baking soda.

Soon you will see that the clear lime water is becoming turbid or cloudy, because the carbon dioxide is reacting with it to form insoluble particles of calcium carbonate. This is a positive test for the presence of carbon dioxide.

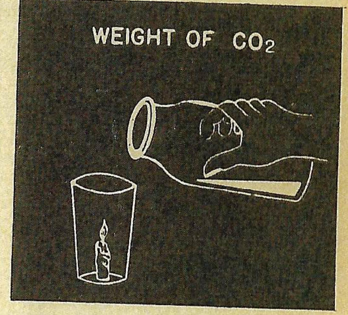
EXPERIMENT 240—Carbon Dioxide Does Not Support Combustion.

Dissolve 6 measures of baking soda in a test tube half full of water. Obtain 2 tumblers or jelly glasses. Make a cardboard or heavy paper cover to fit over the top of one of the glasses and in this cover, make a hole about ¼" in diameter. Have your gas delivery tube and stopper handy. Now add ½ measure of Sodium Bisulfate (No. 7) to the solution of baking soda in the test tube. Quickly insert the gas delivery tube and stopper and place the end of the gas delivery tube through the hole in the piece of cardboard covering one of the tumblers. Thus, the carbon dioxide which is being generated in the test tube is directed into the tumbler. As soon as the effervescence or action in the test tube stops, add another half measure of sodium bisulfate. Again insert the gas delivery tube and stopper and continue to collect the gas in the tumbler. Repeat this until no further effervescence occurs when you add sodium bisulfate. Now strike a match and holding it carefully, lower it into the empty tumbler. If the match is immediately removed, it will not go out, but continues to burn. Now strike another match and holding it the same way, remove the cardboard cover and lower it into the tumbler containing carbon dioxide gas. The match will go out immediately, showing that carbon dioxide does not support combustion.

EXPERIMENT 241—Weight of Carbon Dioxide Gas.

Carbon dioxide is much heavier, that is, much denser, than air. Its high density makes it possible to pour it from one vessel to another almost like a liquid.

Obtain a quart milk bottle and in it place 2 teaspoonfuls of baking soda. Next set a short candle upright in the bottom of a tumbler or jelly glass, using a little melted wax to make it stick to the bottom of the tumbler. The wick should not come higher than half way up the tumbler. Pour about half a glass of vinegar into the milk bottle and while the carbon dioxide gas is being formed and pushing the air out of the bottle, light the candle. Hold the milk bottle over the glass and pour carbon dioxide out of it onto the candle just as if you were pouring a liquid, but be careful not to allow any of the vinegar to pour out. When you have filled the glass up to the level of the flame with carbon dioxide, the candle will go out, for the oxygen supply will be cut off by the heavy CO_2 gas.



Carbon dioxide, because of its relatively high density, tends to collect in abandoned mines and old wells, and so such places should always be tested by lowering a lighted candle into them before anyone descends. Sometimes carbon dioxide is formed by fermentation and accumulation in silos, and death by suffocation may result from entering and remaining in a closed silo.

We have found that carbon dioxide is a heavy gas that will not burn or support combustion. We know that fire will not burn when its supply of oxygen in the air is cut off. Therefore, we reason that carbon dioxide would be a good fire extinguisher. This is true and we find that carbon dioxide is used in many commercial fire extinguishers.

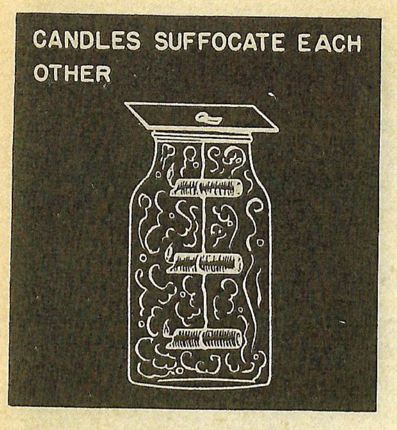
EXPERIMENT 242—Extinguishing a Flame with Carbon Dioxide.

Set a candle in a tumbler so that when lighted the top of the flame will be just below the edge of the tumbler. Now put a spoonful of baking soda in the tumbler and add vinegar, a few drops at a time, being careful to pour the vinegar in so that it will not get on the candle flame. In a few moments the carbon dioxide generated in the bottom of the tumbler will suffocate the candle and the flame will be extinguished.

Some hand fire extinguishers and many chemical fire engines make use of carbon dioxide. They contain a solution of sodium bicarbonate (baking soda) and sulfuric acid, held in separate containers, which are so arranged that when the extinguisher is turned upside down, the chemicals are mixed and carbon dioxide is generated. The pressure of this gas forces a mixture of water and carbon dioxide out of the nozzle in the form of a powerful stream which may be directed upon the fire.

EXPERIMENT 243—Making Candles Suffocate Each Other.

Obtain a quart milk bottle or fruit jar. Now fasten 3 candles to a wire, one above the other, so that when lowered into the bottle, one candle will be near the


CANDLES SUFFOCATE EACH OTHER

bottom, the second about half way up and the third two or three inches below the top.

Light the candles and lower them into the bottle and cover the mouth with a piece of cardboard. As the candles burn, carbon dioxide is formed and since the gas is hot, it rises and collects at the top of the bottle, first extinguishing the upper candle and then those lower down.

EXPERIMENT 244—A Dry Fire Extinguisher.

Place a spoonful of baking soda on a sheet of tissue paper about 4 inches square and wrap the soda in the paper. Put this package on a tin plate and light the paper with a match. Notice that the flame goes out almost at once.

When heated, sodium bicarbonate or baking soda, gives off carbon dioxide just as it does when treated with an acid. The carbon dioxide set free by the heat of the flame promptly settles around the flame and extinguishes it.

EXPERIMENT 245—A Foam Fire Extinguisher.

Pour about half a test tube of thin liquid glue or mucilage into a glass and add 5 or 6 measures of sodium bicarbonate. Stir this mixture thoroughly so that it makes a thin, uniform paste. Now dissolve 4 measures of Sodium Bisulfate (No. 7) in a test tube ½ full of water. Shake the tube vigorously until the solid is all dissolved. Pour the sodium bisulfate solution into the glass containing the paste and a heavy, sticky foam will rapidly form. The bubbles in this foam are filled with carbon dioxide gas and the glue or mucilage makes it so stiff and sticky that it will form a coating on burning wood and smother the fire.

EXPERIMENT 246—Fireproofing Cloth.

Cloth and other inflammable substances may be fireproofed by treating them with chemicals which, when heated, give off vapors that smother the flame.

Dissolve by warming 12 measures of Ammonium Chloride (No. 9) in a test tube ⅓ full of water. Put a piece of cotton cloth, 2 or 3 inches square, in the bottom of the glass and pour the liquid in the test tube over it. Stir the cloth around until it is wet through, then let dry and try to light it with a match. You will find that it will burn while held in the flame, but just as soon as the match flame is removed, it will go out.

Because fabrics treated in this manner cannot be set on fire by sparks or overturned candles, a similar process of fireproofing is sometimes used for treating the curtains and scenery of theaters.

EXPERIMENT 247—Fireproofing Wood.

Sodium silicate solution (water glass) is sometimes used for fire proofing wood, since when it dries, it forms a hard, dense film over the wood through which oxygen cannot penetrate. It is not suitable for cloth or paper because it makes them stiff.

Hold a match by the head and dip about two-thirds of the stick into Sodium Silicate Solution (No. 19). Let the coating dry, and then light the match. There will be no danger of burning your fingers, for the flame will go out as soon as it reaches the film of sodium silicate.

SOLID CARBON DIOXIDE — DRY ICE

Carbon dioxide is a gas under ordinary conditions, but when cooled sufficiently, becomes solid. By compressing the gas, it is possible also to obtain liquid carbon dioxide, but the liquid cannot exist at the pressure of

the atmosphere, and when allowed to escape from a tank under pressure, it evaporates so rapidly that it absorbs enough heat to freeze some of it to the solid state. This is the usual method of making solid carbon dioxide.

Since when the solid absorbs heat it passes directly into the form of gas without first melting to a liquid, it is commonly called dry ice. Large quantities of it are now made for use as a refrigerant in the packing of ice cream and other foods as well as for refrigerator cars. The carbon dioxide used for making it is often obtained as a by-product, for example, from lime kilns or from fermentation vats.

To perform the experiments of this section you will need some dry ice. This may usually be obtained from a soda fountain or ice cream manufacturing plant, but since it evaporates rather rapidly, you had better not get it until you are ready to begin the experiment.

CAUTION: Dry ice can be handled safely with the bare hands so long as it is not squeezed. When pressed against the skin, however, heat is removed so rapidly from the area in contact with the dry ice, that a frostbite is produced which feels very much like a burn. Hence, it is best to handle this material with a pair of pliers or when wearing a leather glove.

EXPERIMENT 248—Testing Dry Ice for Carbon Dioxide.

Stir 2 measures of Calcium Oxide (No. 20) in half a glass of water for a minute or two and then let the undissolved solid settle out. Pour off the clear lime water into another glass, and drop into it a small piece of solid carbon dioxide.

At once the lime water becomes turbid, because of the calcium carbonate formed. This, you will remember, is the usual test for carbon dioxide. If enough carbon dioxide is present, the white precipitate of calcium carbonate may later disappear because an excess of carbon dioxide converts it into soluble calcium bicarbonate.

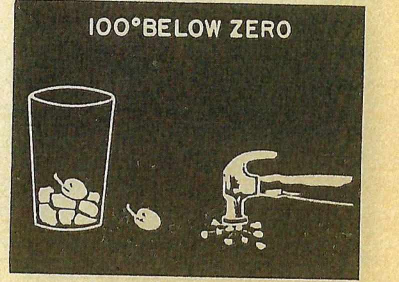
EXPERIMENT 249—How Cold Is Dry Ice?

In a previous experiment we found that although it was easy to freeze water in an ice and salt bath, a strong salt solution had too low a freezing point to solidify in such a cooling bath. We now have, however, a much more powerful cooling agent at our disposal.

Add about 10 measures of table salt to a test tube ½ full of water and shake the tube until the salt is dissolved. Set the test tube in a tumbler containing several pieces of dry ice upon which you have poured some ordinary denatured, or rubbing, alcohol. The alcohol makes better contact with the test tube than the dry ice can, and being itself cooled by the dry ice, it then removes heat rapidly from the test tube. You will soon find that the salt solution has solidified completely.

EXPERIMENT 250—100° below Zero.

Not only salt solution, but many other substances are frozen solid by dry ice. Crush some dry ice in the bottom of a tumbler, using a wooden stick, such as the handle of a hammer. Pour in some ordinary denatured alcohol and place in the resulting mixture various objects, such as a green leaf, a grape, a rubber band, a fragment of raw meat, and a small flower. After several minutes they will be thoroughly chilled. Remove them and examine them quickly before they warm up. You will find that the leaf and the flower have become so brittle that you can crush them into bits. The rubber is very stiff and perhaps even brittle. The grape and the meat are hard like pebbles, and if struck with a hammer while they are still cold, they will crumble into fragments. Of course you must not eat anything which has been soaked in denatured alcohol.


100°BELOW ZERO

The temperature of the alcohol and dry ice mixture is about —75° C. (—103° F.). Cold though it is, dry ice is still hot compared to liquid air, the temperature of which is about —190° C. (—310°F.).

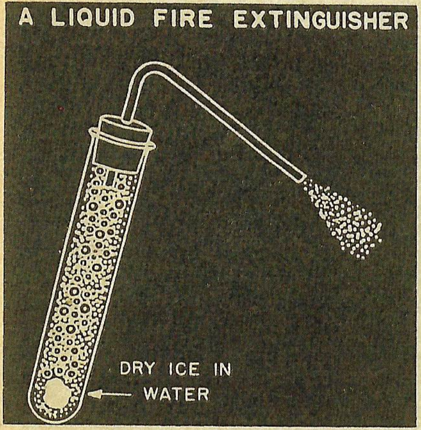
EXPERIMENT 251—Making Soap Bubbles Float on Carbon Dioxide.

Place a few pieces of dry ice in the bottom of a shallow dish, such as a soup plate. In another dish make some soap suds by stirring soap shavings in half a glass of warm water until dissolved. Using a soap bubble pipe, blow some bubbles and allow them to fall into the dish containing the dry ice. If a sufficient amount of the dry ice has evaporated, the dish will be filled with carbon dioxide gas, and since it is considerably more dense than air, the air-filled bubbles will float on it, just as a hollow rubber ball floats on water. Of course you must do this experiment in a room where there are no drafts, since otherwise the carbon dioxide would be blown out of the dish.

EXPERIMENT 252—Extinguishing a Candle with Dry Ice.

Place several pieces of dry ice in the bottom of a ½ gallon jar and then light your candle and place it upright on a tin pan on the table. After about a minute, the carbon dioxide gas formed by the evaporation of the dry ice will have filled the jar. Pick it up and pour the gas quickly over the candle flame, just as if you were pouring water. Being thus flooded with carbon dioxide, the candle can no longer get oxygen, and so the flame is extinguished.

EXPERIMENT 253—Making a Liquid Fire Extinguisher.



Fill a test tube ¾ full of water and drop into it several small lumps of dry ice. Quickly fit the test tube with your gas delivery tube and direct the end of the tube at a can or other convenient receptacle. A strong stream of water will be forced out of the tube by the pressure of the expanding gas. This is the same principle that is employed by the ordinary carbon dioxide fire extinguisher except that in the commercial variety the carbon dioxide is formed by the interaction of sodium bicarbonate and sulfuric acid as we have already learned.

EXPERIMENT 254—Making Carbonic Acid from Dry Ice.

Fill a tumbler half full of water, and add a very small quantity of Sodium Carbonate (No. 4) (⅛ measure should be enough). Now add 2 drops of Phenolphthalein Solution (No. 22). The liquid should be a pink color. Drop in several small pieces of dry ice, and as they evaporate and the gas dissolves in the water, the pink color of the phenolphthalein will fade away, showing that the originally basic solution has now become neutral and perhaps acid. Test it with blue Litmus Paper (No. 18) and you will find that the paper turns red, showing that the solution is acid. In the following section we are going to learn some more about the acid properties of carbon dioxide.

CARBONATES

Since carbon is a non-metallic element its oxide, carbon dioxide, forms an acid when combined with water. This acid is called carbonic acid, and it is made very simply by dissolving carbon dioxide in water.

$$CO_2 + H_2O \rightleftarrows H_2CO_3$$

We write the reaction with a double arrow because carbonic acid decomposes very easily, and in any solution of carbon dioxide, molecules of carbonic acid are continually forming and breaking down, so that there is a balance or equilibrium, between the forward and the reverse reactions.

Although carbonic acid is an extremely weak acid, that is, it produces only a few hydrogen ions, it is able to turn blue litmus paper red and to combine with many bases to form salts.

EXPERIMENT 255—Carbonic Acid.

Into a test tube half full of water drop a very small strip of red Litmus Paper (No. 18) and also a strip of blue Litmus Paper (No. 18).

Put about half a teaspoonful of baking soda or powdered washing soda into another test tube, add a small quantity of water and a few drops of vinegar. Immediately insert the gas delivery tube and stopper into this test tube and then put the end of the gas delivery tube in the test tube containing the 2 pieces of litmus paper. Carbon dioxide gas will be given off from the soda and vinegar and will pass through the delivery tube and bubble into the water containing the litmus paper. Keep a flow of gas for several minutes, occasionally adding 2 or 3 drops of vinegar to the test tube that contains the soda. Soon you will notice that the blue litmus paper is changing over to the same color as the red litmus paper, indicating that the carbon dioxide gas has dissolved in the water to form carbonic acid.

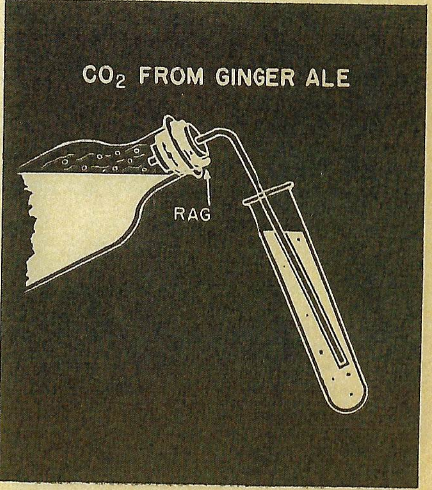
When the change from blue to red is completed, warm the test tube of carbonic acid gently and notice the large number of gas bubbles which form on the sides of the tube. Carbon dioxide is much less soluble in water at higher temperatures, so that some of it must escape as the solution is warmed. Boil the liquid for a few moments, taking care that it does not spurt out of the tube. Boiling decomposes the carbonic acid and drives carbon dioxide out of the solution, so that now if you drop in a piece of blue litmus paper, it will remain blue, showing that an acid is no longer present.

The color changes in this experiment will depend a good deal upon the impurities, such as hardness, in the water which you use. Rain water or distilled water will work best.

Carbonated beverages, such as gingerale and soda water, are saturated with carbon dioxide under pressure. When the pressure is released, so much gas cannot remain dissolved, and so it comes out in the form of bubbles. Soda water is so-called because it was formerly made with carbon dioxide that had been prepared from soda.

EXPERIMENT 256—Identifying the Gas from Gingerale.

Wrap a narrow strip of cloth or paper around the cork of your gas delivery tube so that it will fit the mouth of a bottle of gingerale or other carbonated beverage. Fit your delivery tube into the mouth of a newly opened bottle and place the other end of the tube in a test tube ¾ full of clear lime water, made by adding half a measure of Calcium Oxide (No. 20) to a test tube ¾ full of water, shaking thoroughly, allowing to settle, and then pouring off the clear lime water into another test tube. Allow the gas to bubble through the lime water for a few minutes. The white precipitate of calcium carbonate proves that the gas is carbon dioxide.



The carbonates of many metals are insoluble in water, so they can be precipitated from solutions by means of a soluble carbonate, such as sodium carbonate. Since carbonic acid is a very weak acid, solutions of soluble carbonates are usually alkaline and turn red litmus paper blue.

EXPERIMENT 257—A Test for Carbonates.

Suppose you have some substance which you believe is a carbonate. How can you test it? Since all carbonates give off carbon dioxide gas when treated with acid, you need only to add acid to your unknown substance and test it with lime water to see whether the gas, if any, which comes off is carbon dioxide.

Prepare some clear lime water by stirring 3 measures of Calcium Oxide (No. 20) in a glass half full of water for a minute or two. Let the undissolved solid settle to the bottom and fill a clean test tube ¼ full of the clear liquid.

In another test tube place 4 measures of the substance you want to test. It should be crushed or ground to small pieces. Add to it 5 measures of Sodium Bisulfate (No. 7) and a few drops of water, and then fit the test tube quickly with the gas delivery tube and stopper. Dip the other end of the gas delivery tube into the lime water in the other test tube. If carbon dioxide gas is being set free, it will make the lime water cloudy, and prove that your unknown substance contains a carbonate.

Test in this manner any of the following substances:

Baking soda (sodium bicarbonate), washing soda (sodium carbonate), baking powder (a mixture of sodium bicarbonate with potassium acid tartrate, calcium acid phosphate or alum), oyster shells (which consist chiefly of calcium carbonate), chalk, tooth powder, and plaster (which also contains calcium carbonate).

EXPERIMENT 258—Making Sodium Carbonate from Sodium Bicarbonate.

Dissolve 1 measure of baking soda (sodium bicarbonate) in a test tube half full of water and add 1 or 2 drops of Phenolphthalein Solution (No. 22). Unless your water supply is quite alkaline, the solution will be colorless.

Now warm the test tube over a flame and notice how much carbon dioxide comes off in the form of bubbles even before the liquid boils. Boil carefully for a few moments, and the solution will slowly turn pink, showing that it has become alkaline due to the formation of sodium carbonate.

$$\underset{\text{(sodium bicarbonate)}}{2\ NaHCO_3} + \text{Heat} \rightarrow \underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O}$$

EXPERIMENT 259—Making Sodium Bicarbonate from Sodium Carbonate.

The reverse of the preceding experiment is to add enough carbon dioxide to sodium carbonate to convert it into the bicarbonate. The equation for the change is just the reverse of the one we wrote above.

Dissolve ¼ measure of Sodium Carbonate (No. 4) in a test tube half full of water and add 1 drop of Phenolphthalein Solution (No. 22), coloring the solution red.

Now generate CO_2 gas with sodium bicarbonate and vinegar and pass carbon dioxide through the sodium carbonate solution for several minutes. Slowly the pink color of the phenolphthalein begins to fade as the sodium carbonate absorbs carbon dioxide and changes into sodium bicarbonate. The final loss of color comes rather suddenly.

$$Na_2CO_3 + CO_2 + H_2O \rightarrow 2\ NaHCO_3$$

EXPERIMENT 260—Decomposition of Dry Sodium Bicarbonate.

Fill a test tube half full of water and add 1 drop of Phenolphthalein Solution (No. 22). Unless your water supply is too alkaline, there will be no color. Drop in ½ measure of baking soda and notice that the solution remains colorless.

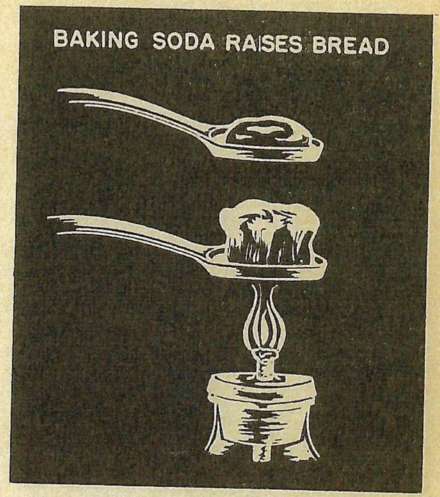
Now place several measures of baking soda in your spoon and heat it strongly for several minutes over a flame. Allow the substance to cool and then drop it into the water containing phenolphthalein. Shake the test tube and the mixture will turn pink, showing that some of the dry sodium bicarbonate has been converted to sodium carbonate.

EXPERIMENT 261—How Baking Soda Raises Bread.

Ordinary baking soda is simply sodium bicarbonate, and as you saw in the preceding experiment it gives off carbon dioxide gas when heated. Therefore, when

it is mixed into biscuit or cake dough, it forms gas bubbles throughout the dough during baking, producing a porous light texture. Weakly acid substances, such as sour milk or cream of tartar, are often added to assist in the liberation of carbon dioxide.

Put a teaspoonful of flour into a cup and add 4 measures of baking soda. Stir water into the flour in small portions until you have a good smooth dough. Place a piece of this dough about the size of a marble in your spoon and heat it gently over the alcohol lamp flame. Notice how it puffs up and becomes porous due to the gas which forms inside.


BAKING SODA RAISES BREAD

When lime is used in mortar, it slowly absorbs carbon dioxide from the air and is converted into calcium carbonate, just as carbon dioxide forms a precipitate in lime water. Calcium carbonate is the chief constituent of limestone and marble, and when it is formed from the lime in the mortar, it gives strength to the masonry. If we expose some lime water to the air, calcium carbonate will soon be formed.

EXPERIMENT 262—Limestone Ice.

Make some clear lime water by stirring 2 measures of Calcium Oxide (No. 20) in half a glass of water for a minute or two. Let the suspended solid settle to the bottom and pour off the clear solution into a saucer. Let it stand in a quiet place overnight, and in the morning you will find that a thin white crust has formed over the lime water. Although only 0.04% of the air is carbon dioxide, the calcium hydroxide solution has absorbed enough of it to make this film of carbonate. Add a drop or two of vinegar and watch the crust dissolve, giving its carbon dioxide back to the air.

EXPERIMENT 263—Strontium Carbonate.

Dissolve 1 measure of Strontium Chloride (No. 10) in a test tube half full of water and add a few drops of sodium carbonate solution prepared by dissolving 3 measures of Sodium Carbonate (No. 4) in a test tube half full of water. The white precipitate is strontium carbonate. Strontium is an element which closely resembles calcium in its properties, so it is not surprising to find that their carbonates are very similar. Strontium carbonate is found in nature as the mineral strontianite.

The carbonates of some metals are hydrolyzed to so great an extent that they cannot exist in contact with water. When you try to make them by the following method you obtain a precipitate which is a sort of cross between the carbonate of the metal and the base, or hydroxide, of the metal, that is, you obtain a basic carbonate. The exact composition of such a compound often depends upon the temperature and concentration of the solutions from which it is precipitated.

EXPERIMENT 264—Basic Magnesium Carbonate.

Add 1 measure of Epsom salts (magnesium sulfate) to a test tube half full of water and dissolve the solid by shaking the tube. Now add a few drops of sodium carbonate solution prepared by dissolving 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water, and a white precipitate of basic magnesium carbonate will be formed. This is somewhat similar to the mineral magnesite.

EXPERIMENT 265—Basic Cobalt Carbonate.

Repeat the preceding experiment using about 8 drops of Cobalt Chloride Solution (No. 27) in place of the magnesium sulfate. When you add sodium carbonate solution, a light blue precipitate of basic cobalt carbonate is formed.

EXPERIMENT 266—Basic Copper Carbonate.

Place 1 measure of Azurite (No. 39) and 3 measures of Sodium Bisulfate (No. 7) in a test tube. Fill the tube half full of water and shake it for a few moments until the solution becomes clear. Now add a few drops of sodium carbonate solution prepared as before, and a very pretty blue precipitate of basic copper carbonate will be formed. The minerals malachite and azurite are basic carbonates of copper.

EXPERIMENT 267—Basic Zinc Carbonate.

Put a small piece of zinc about the size of the end of the measure into a test tube and add 2 measures of Sodium Bisulfate (No. 7). Fill the test tube ¼ full of water and warm it gently over a flame for a minute or two. Pour off the clear solution into another test tube, and add to it a few drops of sodium carbonate solution prepared as in the preceding experiments. The white precipitate is basic zinc carbonate.

A few metals refuse to form even basic carbonates, and when sodium carbonate is added to solutions of their salts, the base, or hydroxide, of the metal is formed instead of the carbonate or basic carbonate.

EXPERIMENT 268—Precipitation of Ferric Hydroxide by Sodium Carbonate.

Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube half full of water, and add a little sodium carbonate as in the preceding experiment. A dense brown precipitate forms at once and carbon dioxide gas is given off. It is just as if ferric carbonate first formed but then immediately hydrolyzed, giving ferric hydroxide and carbonic acid.

EXPERIMENT 269—Precipitation of Aluminum Hydroxide by Sodium Carbonate.

Dissolve 1 measure of Aluminum Sulfate (No. 12) in a test tube half full of water, add a little sodium carbonate as in the preceding experiment. A dense white precipitate of aluminum forms at once and carbon dioxide gas is given off. Aluminum carbonate, like ferric carbonate, immediately hydrolyzes, giving aluminum hydroxide and carbonic acid.

FUELS

If mankind had never learned how to use fuel, there would be no civilized human race. Long before the beginning of history, our primitive ancestors were using fire to keep them warm and to cook their food. Today we are still using fire for those two purposes, but we have learned to use it also in many other ways. Fuel burned under boilers gives us steam to turn the wheels of industry, to drive ships across the sea, and to generate electricity, the bearer of light and power. Gasoline and fuel oil burned in the cylinders of engines propels us along the highways, the railways and through the air. In blast furnaces and smelters, it is the burning of fuel which unlocks metals from the ores in which they are hidden.

The most important fuels of our modern times are coal, petroleum and natural gas. All of these fuels have been formed from the remains of prehistoric vegetation, which became covered with millions of tons of earth and rock. During the ages, the heat and pressure gradually converted this buried vegetation into carbon (coal) and hydrocarbons (petroleum and natural gas).

Let us begin the study of fuels by seeing what we can learn from a candle. Candles are usually made from paraffin, a substance obtained from petroleum, and generally have some stearic acid added to harden them. Paraffin is a mixture of compounds composed of the elements carbon and hydrogen. Such compounds made up only of carbon and hydrogen are called hydrocarbons. When a hydrocarbon burns completely, the carbon is converted into carbon dioxide, and the hydrogen into water. We learned, when we studied water, that when a candle burns, the hydrogen contained in the candle combines with the oxygen of the air, forming water. Now let us prove that carbon dioxide is also formed when a candle burns.

EXPERIMENT 270—Showing that Carbon Dioxide Is Formed When a Candle Burns.

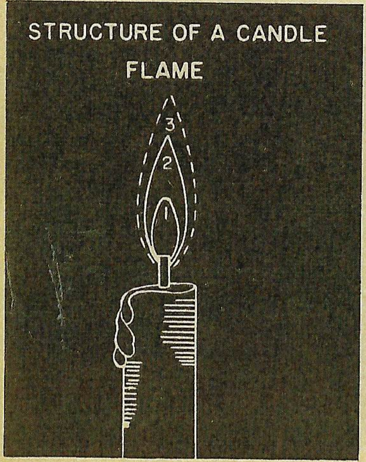
Place 2 measures of Calcium Oxide (No. 20) in half a glass of water and stir the mixture for a minute or two. Let the solution stand until the powder has settled to the bottom and then pour off the clear liquid into another glass. This liquid is lime water.

Set a candle in the middle of an ordinary sauce or baking pan, and stick it tight with a few drops of molten wax. Pour about 2 inches of water into the pan, light the candle and set an inverted bottle or fruit jar over it.

When the candle flame has gone out, remove the jar or bottle from the water, turning it right side up as soon as possible to keep the heavy gas from spilling out, and quickly pour in a little of the lime water. Put your hand over the mouth of the bottle or jar and shake it for a moment. If you will look closely you will see that the lime water is now turbid with white particles. As we have already learned, the formation of this white precipitate of calcium carbonate is a test for carbon dioxide.

A flame is simply a body of burning gas. In order to burn with a flame, a substance must first be vaporized, or converted into a gas. A candle burns with a flame because the wax, melted by the heat, climbs up the wick, and is converted into vapor, which then combines with the oxygen of the air so rapidly that it gives off heat and light.

EXPERIMENT 271—The Structure of a Candle Flame.



Look closely at a candle flame and notice its two main parts. First, the dark zone around the wick, and second, the bright yellow luminous part of the flame. Outside this is a third zone, consisting of the invisible, hot products of combustion.

The dark inner zone consists of unburned gas given off from the wick of the candle. Since there is no air at the center of the flame, this gas cannot burn at first, and so this portion of the flame is relatively cool, as you can show by thrusting a match stick into the center of the flame just above the wick. The part of the stick held in the dark zone does not char as soon as the part passing through the side of the flame.

Candle wicks are woven in such a way that as the candle burns, the end of the wick bends over into the hot outer part of the flame. This keeps it burned off short and prevents the flame from smoking.

EXPERIMENT 272—Removing Gas from a Candle Flame.

With care, a portion of the unburned gas can be drawn off from the center of the flame. Hold one end of a glass tube in the flame directly over the wick. Let the tube slant upwards so that the hot gas will rise through it and issue from the other end, where it can be lighted with a match.

EXPERIMENT 273—Igniting Candle Gas.

Blow out a candle with a long wick that is burning and after an instant relight the smoke at a little distance from the end of the wick. Notice how much easier it is to light a candle just after it has been extinguished than after it becomes cold. Gas is still coming off from the hot wick and is ready to be ignited, but in lighting a cold candle, wax must be melted and vaporized before it can burn. In an alcohol or kerosene lamp, as well as in a candle, the fuel must be vaporized before it can burn.

EXPERIMENT 274—The Luminous Zone—Making Lampblack.

The second, or bright yellow part, of the flame contains particles of white-hot carbon, formed by the action of heat on the gas of the inner zone. These particles burn completely in the outside of the flame. Their presence can be proved by holding a cold spoon or piece of glass in the flame for a few moments. It will become covered with a black deposit of soot, or lampblack. The cold object chills the flame and prevents the carbon from being completely burned.

In making lampblack on a large scale, natural gas is burned in large chambers and the flame is chilled by iron pipes through which water circulates. Large quantities of lampblack are used in printing ink and as a toughening agent in rubber.

WOOD

Wood is still used as a fuel in some localities. This important material contains carbon, hydrogen, oxygen and nitrogen in the form of complex compounds. When wood is burned with a plentiful supply of air, we would naturally expect to form carbon dioxide and water with smaller amounts of other substances, which are found largely in the ash or noncombustible part of the wood.

EXPERIMENT 275—Testing for Carbon Dioxide in the Combustion Products of Wood.

Light a splinter of wood, somewhat larger than a match stick, and when it is burning brightly drop it into a milk bottle or fruit jar, and cover the top of the jar with a piece of cardboard.

When the flame has gone out, pour some clear lime water into the jar and shake it around. In this case, you will find that the lime water turns white or turbid, showing that carbon dioxide is formed when wood burns.

EXPERIMENT 276—Testing for Carbon Dioxide in the Combustion Products of Paper.

Since most paper is composed largely of wood fibers, we would expect to find the same products of combustion from burning paper as we do from burning wood. Light a small strip of paper and drop it into a jar or milk bottle, covering the top of the jar with a piece of cardboard. When the flame has gone out, test for carbon dioxide gas by pouring in a little clear lime water and shaking the jar.

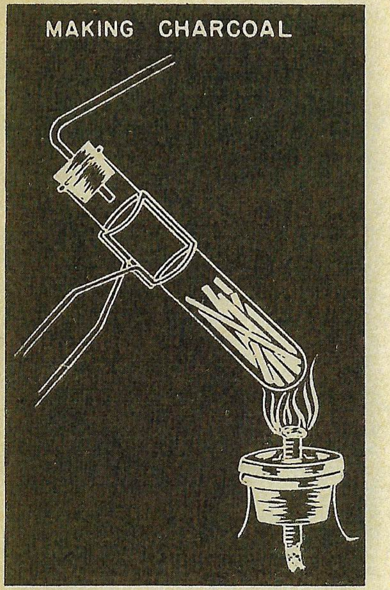
EXPERIMENT 277—Showing that Water Is Formed When Wood Burns.

Light a small splint of wood and as soon as it is burning well, hold it upright and invert a dry, cool glass tumbler over the flame. After a few seconds remove the tumbler and notice that moisture is formed on the inside surface of the glass. This proves that water is formed when wood burns.

EXPERIMENT 278—Showing that Water is Formed When Paper Burns.

Roll a piece of paper, about 4 inches square, into a loose tube. Light one end of this tube and as soon as the paper is burning well, hold it in an upright position and invert a dry, cool glass tumbler over the flame. After a few seconds remove the tumbler and again examine it for moisture on the inside surface.

We have learned that when wood is burned with a plentiful supply of air (oxygen), carbon dioxide and water are formed. When wood is heated, however, in a closed vessel or retort so that the oxygen of the air cannot unite with it, a complicated chemical decomposition occurs and a variety of compounds is obtained. A combustible mixture of gases are given off from the wood, also vapors of various liquids, wood alcohol, acetic acid (wood vinegar), some tar-like substances and moisture. The residue is charcoal, composed almost entirely of carbon, along with the mineral matter (ash) of the original wood.



EXPERIMENT 279—Making Charcoal.

Break the heads off 5 or 6 wooden matches and put the sticks in the bottom of a test tube. Now heat the tube, using your test tube holder, over an alcohol lamp or gas flame.

When smoke begins to pass from the test tube, hold a piece of moistened blue Litmus Paper (No. 18) in the smoke. Notice that it is turned red, showing that there is acid in the gas. Now fit the tube with your gas delivery tube and continue heating. You will find that the gas which issues from the tube is inflammable, and after the distillation has continued for a few minutes you can light it as it comes out of the tube.

Continue heating until gas no longer comes off, then let the test tube cool and empty out the charcoal which remains, using your stirring rod to loosen it from the sides, if necessary. Charcoal is an impure variety of amorphous carbon, that is, carbon which is not crystalline.

COAL

When coal is heated in a closed vessel, coal gas is evolved and also ammonia, along with a great variety of oily and tar-like liquids. The distillation or carbonization of coal is an important source of ammonia for use in fertilizers, and from the oil and the coal tar may be prepared the raw materials of dyes, drugs, disinfectants, high explosives, perfumes, and a multitude of other chemical products. The residue left behind in the retorts is coke, composed, like charcoal, of amorphous carbon. Its chief use is in blast furnaces and smelters for obtaining metals from their ores.

Although a large fraction of our coal production is heated in coke ovens to obtain coke along with the valuable by-products we have mentioned, most of it is still burned directly, wasting millions of tons of ammonia and precious organic chemicals. Coke is clean to handle and smokeless to burn, and in the future coal will probably be more and more converted into coke before being burned, the by-products—gas, ammonia, oils, and tar—being collected and used in industry.

EXPERIMENT 280—Making Coke.

Put 4 measures of powdered soft coal into a test tube and heat it over a flame. Soon you will notice fumes coming from the mouth of the test tube. Hold a piece of cold glass or porcelain at the mouth of the tube as you heat it, and a yellow or brown liquid will condense upon it. This is similar to crude coal tar. Fit your gas delivery tube to the test tube, and when gas begins to issue from it, light it and it will burn. When purified, this gas is the same as the coal gas which is often used for heating.

When smoke no longer issues from the coal, stop heating, and after the tube cools, examine the substance which remains. You will find that it is coke.

OIL FROM THE EARTH

Petroleum is an oily liquid found below the surface of the earth at many places. Usually it is a dark, thick liquid, often with an unpleasant odor, but different samples of petroleum vary greatly in their appearance and properties. Sometimes it is almost as clear and fluid as gasoline, while other specimens may be black and semi-solid, like asphalt. It is usually found in strata of porous rock or sand, sometimes near the surface, in other places miles down in the earth.

Petroleum consists principally of a large assortment of hydrocarbon compounds, with a small percentage of nitrogen and sulfur compounds. Samples coming from different places may have very different composition, and not all are equally suited for all purposes. Most petroleum is supposed to have been formed by the decomposition of plant and animal remains which long ago were buried deep in the earth.

In refining petroleum, it is boiled in a still, and the vapors are led through a high tower in which they are condensed back to liquid. The heavy oil with a high boiling point condenses first near the bottom of the tower; then higher up, the lighter oils and kerosene; then naphtha, used as a solvent for rubber and for lacquers; and finally, the light, volatile gasoline, having the lowest boiling point of all. Each of these products must be further purified before it is ready for market. The gasoline and kerosene are washed with alkali and sulfuric acid to remove sulfur compounds and other undesirable substances. The oils are chilled to —40° F., at which temperature paraffin wax freezes out as a solid. From some of the heavier oils, vaseline, or petrolatum, is separated by a similar procedure. Petroleum from some wells, especially in California and Texas, leaves a residue of asphalt in the still after the volatile constituents have boiled off.

Since 1920 the rapid growth in the number of automobiles has created an ever increasing demand for gasoline, a demand which could not possibly be satisfied by the natural gasoline in petroleum. A large amount of our gasoline at present is made from the heavier, less volatile portions of the petroleum by a process called *cracking*. Kerosene and various grades of oil are heated to about 950° F. in a coil of pipe under pressure of approximately 400 pounds per square inch. The vaporized oil passes into a huge steel chamber, where the large hydrocarbon molecules crack apart into smaller fragments, more suitable for motor fuel. Cracked gasoline contains a different type of hydrocarbon compound from that which makes up natural gasoline, and this makes it burn more smoothly without knocking in modern high-compression engines. Cracking, therefore, produces not only more gasoline from a given amount of petroleum, but also better gasoline.

Further improvement in the quality of gasoline results from adding small amounts of certain compounds to it. The most commonly used is lead tetraethyl, $Pb(C_2H_5)_4$, which has a very marked effect in reducing knocking in high-compression engines.

Diesel engines are now used to pull many trains and for many passenger ships, and smaller models are being developed for automobiles and airplanes. The big advantage of Diesel engines is that instead of gasoline, they use heavy, comparatively non-volatile oil. They are, therefore, much cheaper to operate than ordinary gasoline engines. Since their fuel cannot be vaporized in a carburetor and sucked into the cylinder, it is sprayed in under pressure. The hot, compressed gases in the cylinder cause it to ignite, so that no spark plug is needed.

One of the things which oil chemists and engineers want to know about oils is the temperature at which they will take fire. If this temperature, called the flash point, is too low, there is danger of fire or explosion when the oil is used to lubricate machinery. The flash point of kerosene is important when it is used in a lamp or a stove.

To determine the flash point of an oil, it is heated slowly in a metal cup, and at intervals, a small flame is passed across it at a certain height above its surface. Its temperature is read on a thermometer when it first takes fire with a flash as the flame is passed over it.

EXPERIMENT 281—Flash Point of an Oil.

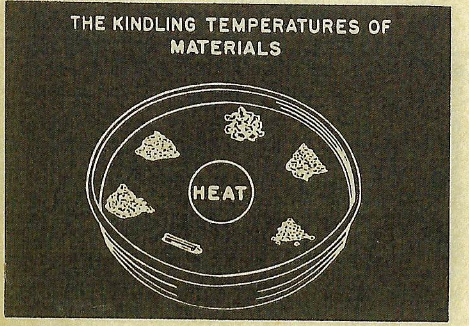
Pour a few drops of lubricating oil into your spoon and warm it slowly over a flame. Soon fumes will begin to be evolved. Test them from time to time by bringing the edge of the spoon into the flame. When the flash point of the oil is reached, the fumes will take fire. Try this experiment with automobile crankcase oil, with light machine oil, and with a few drops of kerosene, and notice the differences among them.

EXPERIMENT 282—Flash Point of Alcohol.

Put a few drops of alcohol which you use for your lamp into the spoon and heat it gently over a flame. Notice how much sooner it takes fire than the oil in the preceding experiment. Because of their low flash point, rubbing alcohol and radiator alcohol should be used with care when open flames are near.

Just as oils and other inflammable liquids catch fire at different temperatures, so solid substances differ greatly in the ease with which they can be set on fire. The temperature to which a substance must be heated before it begins to burn is called its kindling point, or ignition temperature.

EXPERIMENT 283—Ignition Temperature.


THE KINDLING TEMPERATURES OF MATERIALS
HEAT

Obtain a tin plate and arrange in a circle, about 2 inches in diameter, all of the following substances which you can obtain. Put each substance in a little pile, so that there will be about an inch between piles. Use 2 measures of powdered coal, 2 measures of powdered charcoal, a small piece of paper crumpled up into a ball, a short piece of match stick, the head of a safety match, 2 measures of Sulfur (No. 1) and 2 measures of sugar.

Put the pan on a support and set your alcohol lamp under the center so that the flame will strike the pan exactly in the center of the circle, an equal distance from each material.

Watch closely and see which substance catches fire first, and which is the last to burn.

EXPERIMENT 284—Lowering the Ignition Point of Sugar.

Holding a lump of sugar in your test tube holder, bring one end of it into a candle flame and try to light it. Although the sugar turns black and melts, you will find it very difficult to make it burn with a flame.

Now touch the end of the lump to some cigarette ashes so that a small amount of the ash sticks to the sugar. Again hold the sugar in the flame, and this time you will find that it burns easily and that it will continue to burn even after you remove it from the flame. The ash contains small quantities of metallic oxides which act as a catalyst and promote the combustion of the sugar. Since the tobacco ash does not itself burn but nevertheless promotes the combustion of the sugar, it is an example of a catalyst. A catalyst is a substance which increases the speed of a chemical reaction without itself being permanently changed. Instead of cigarette ash, numerous other substances, such as iron rust or sodium carbonate, may be used in this experiment.

MATCHES

The world uses a greater number of matches than of any other manufactured article, yet the cheap, convenient, and safe modern match is a rather recent invention. The history of man's methods of making fire, from the cave man rubbing two sticks together down to the present scratching of a match, is extremely interesting, and many ingenious methods have been invented, used for a while and then discarded in favor of something better.

Modern matches are made by automatic machines which turn them out at the rate of millions a day. In the case of wood matches, an automatic cutter slices the blocks and inserts the match sticks into holes in an endless belt on which they pass through the dipping machine. In making paper matches, the cardboard is diecut and long strips are fed through the dipping machine, after which they are cut to correct sizes.

The strike-anywhere match contains phosphorus sesquisulfide, P_4S_3. The head of a safety match contains no phosphorus, but only a mixture of an oxidizing agent with an easily combustible material. The striking panel on the side of the box or on the folder which holds the matches, however, is coated with a mixture which contains red phosphorus and friction upon this panel ignites the head of the match.

EXPERIMENT 285—Proving that Matches Contain Sulfur.

Cut off the heads of several matches. Place these in a test tube and add 3 or 4 drops of water, and upon shaking the tube, a thin paste will be formed. Place a drop or two of this paste on a shiny piece of silverware or a silver coin. Heat the silver over your alcohol lamp until the pasty material ignites. After this paste burns, notice the black stain which appears on the silver. This is due to the formation of silver sulfide, formed by the sulfur in the matches uniting with the silver. The formation of black silver sulfide proves that the matches contain sulfur.

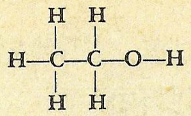
ORGANIC CHEMISTRY

Many years ago chemists noticed that substances obtained from animal or vegetable life always contain carbon. Substances obtained from minerals on the other hand usually did not contain carbon, with the outstanding exception of the carbonates of the various metals. Based on these observations, the term organic chemistry was applied to compounds derived from substances of animal or vegetable origin. The term inorganic chemistry was applied to compounds derived from substances of mineral origin.

The fact that all substances derived from animal and vegetable life contain carbon led also to the belief that the carbon atom and life go hand in hand. Organic chemistry was, therefore, defined as pertaining to substances which contain carbon and are obtained from life. Later it was found that organic substances containing carbon could also be made in the laboratory without the aid of life. Hence, today we define organic chemistry as the chemistry of carbon compounds.

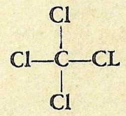
In inorganic chemistry, a majority of the reactions are based on ionization and these reactions occur as a rule almost instantaneously when the substances are brought together in the form of water solutions. In organic chemistry we find that most organic compounds are not soluble in water and are not readily ionized. As a rule organic reactions are slower and more frequently require the aid of heat as well as time.

The organic chemists are principally interested in the arrangement of atoms within the molecule, and this accounts for the practice of showing molecular structure in detail when writing organic formulas and equations. For example, the formula of ethyl alcohol, or grain alcohol, might be written simply as C_2H_6O. It is more informative, however, to show how the atoms are arranged in the molecule by writing a structural formula:



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & — & C & — & C & — & O—H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

Such a detailed structural formula is often abbreviated to CH_3CH_2OH or even to C_2H_5OH, at the sacrifice, however, of some clarity, at least for the beginner. Methanol, or methyl alcohol, has the formula CH_3OH. See if you can figure out how the complete structural formula of this compound would look. Carbon tetrachloride is CCl_4. The structural formula, therefore, would be written:



$$\begin{array}{ccccc} & & Cl & & \\ & & | & & \\ Cl & — & C & — & CL \\ & & | & & \\ & & Cl & & \end{array}$$

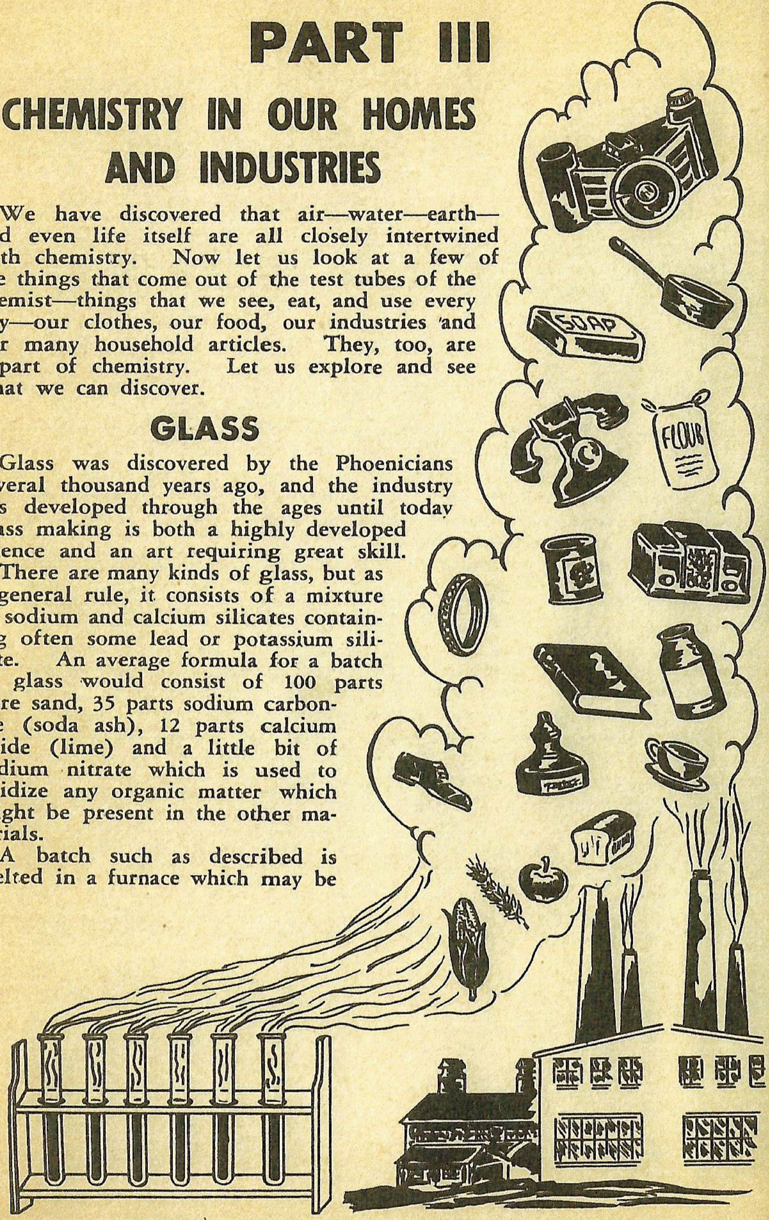
The writing of organic formulas on paper can be likened to the drawing of an architect's plans for a building. Today the organic chemist can even predict in advance the probable properties of new organic compounds before they are made in the laboratory.

Almost all modern dyes, drugs, and explosives are organic compounds. A few familiar examples are indigo, sulfanilamide, procaine, thiamine, ascorbic acid, penicillin, TNT, and nitroglycerine. Foods, including fats, carbohydrates, proteins, vitamins, and flavors, are organic; rubber, natural and synthetic, resins, plastics, fuels, paints, paper, wood, and clothing are composed of organic substances. The life of all plants and animals is absolutely dependent upon reactions of organic compounds. Thus we see that all of us come in contact with organic chemistry a thousand times a day.

In the following part of this book, "Chemistry in Our Homes and Industries," you will take up several practical applications of chemistry which are ordinarily classified as organic chemistry.

PART III

CHEMISTRY IN OUR HOMES AND INDUSTRIES



We have discovered that air—water—earth—and even life itself are all closely intertwined with chemistry. Now let us look at a few of the things that come out of the test tubes of the chemist—things that we see, eat, and use every day—our clothes, our food, our industries and our many household articles. They, too, are a part of chemistry. Let us explore and see what we can discover.

GLASS

Glass was discovered by the Phoenicians several thousand years ago, and the industry has developed through the ages until today glass making is both a highly developed science and an art requiring great skill.

There are many kinds of glass, but as a general rule, it consists of a mixture of sodium and calcium silicates containing often some lead or potassium silicate. An average formula for a batch of glass would consist of 100 parts pure sand, 35 parts sodium carbonate (soda ash), 12 parts calcium oxide (lime) and a little bit of sodium nitrate which is used to oxidize any organic matter which might be present in the other materials.

A batch such as described is melted in a furnace which may be either a clay pot that is set in a hot furnace, or a tank furnace, which is merely a tank built directly in a furnace made of heat-resistant bricks. The batch of material is left in the hot furnace for 24 hours or more until it melts and becomes fluid and free from bubbles. This melted glass is then ready to work and form into the desired shapes.

EXPERIMENT 286—Making Soft Glass.

Put 5 drops of Sodium Silicate Solution (No. 19) in your spoon and mix thoroughly with 1 measure of Sodium Carbonate (No. 4) and $\frac{1}{3}$ measure of Calcium Oxide (No. 20). Hold your spoon in the flame of your alcohol lamp. You will notice that when you heat the mixture it will swell, turn white, froth, and finally assume a shiny glass-like appearance. This is soft glass, which is a soda-lime silicate. If you find that the heat of your alcohol lamp flame is not sufficient to melt the glass, hold your spoon over the flame of a gas burner.

To make plate glass, the molten glass is poured on a large, flat table and rolled to the desired thickness by a metal roller. The plate glass is then put in an oven and annealed, or allowed to cool very slowly. It is then smoothed and polished.

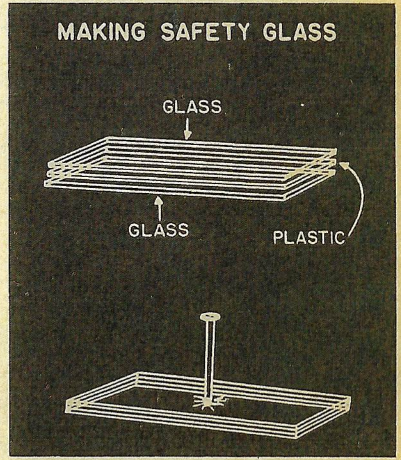
Wire glass is plate glass into which wire has been imbedded during the rolling process. Safety glass which will not shatter is made by cementing a sheet of transparent plastic between two pieces of plate glass.

EXPERIMENT 287—Making Safety Glass.

You can make a piece of safety glass from two microscope slides. Cover one side of each with Canada balsam or cellulose acetate cement, such as is used in gluing model airplanes. Then cut a piece of clear plastic and place this between the coated slides. Press the two pieces of glass and the plastic together and place them in a clamp to dry. In two or three days, after the cement has dried, you will notice that you can see through the two pieces of glass and plastic; just as well as if you were looking through only one piece of glass.

EXPERIMENT 288—Testing Safety Glass.

Test the piece of "safety" glass which you made in the preceding experiment by driving a nail through it and note that it does not shatter, but holds together and merely cracks. Keep your face at a distance as you test the glass.


MAKING SAFETY GLASS

Window glass is now made by machinery but formerly it was made entirely by hand. The workman would gather a lump of molten glass on the end of a hollow iron pipe and by alternately blowing and swinging he produced a large hollow cylinder of glass. The cylinder was allowed to harden and then was cut lengthways along one side. The next step consisted of putting this cut cylinder into an oven and gradually heating it until the glass began to soften and the cylinder flattened out into a pane. Machine-made window glass is made very much the same way, except that the operation of dipping the hollow pipe into the molten glass and blowing are done mechanically.

Bottles and similar heavy glassware are made by dropping some molten glass into a mold and forcing the glass to take the shape of the mold by means of a plunger or compressed air, or both. On modern bottle making machines each operation is automatic and bottles are turned out at an amazing rate of speed.

Almost without exception, articles made of glass have to be cooled very slowly, or annealed, as the process is called. If glass is cooled suddenly, the outside will harden while the interior of the glass is still in a semi-liquid state and under these conditions great strains develop so that the glass is likely to fall to pieces due to a change in temperature or a scratch on the surface. To obtain glass free from strain, the finished article, while it is still hot, is placed in a heated room or oven and allowed to cool very slowly. The time of cooling or annealing varies from several hours to a week or more, depending on the thickness and composition of the glass.

An interesting recent development in the glass industry is a strong comparatively flexible glass which is very resistant to sudden temperature changes and to shock. A piece of it ¼ inch thick will not crack when molten lead is poured over it while it rests on a block of ice. When it is broken by a blow, the fragments are rounded instead of sharp and jagged. It is made by a carefully controlled chilling of the glass from a high temperature in such a way that the interior contracts, compressing the surface layers.

Many special kinds of glass are made; for example, cut glass, which has a brilliant, shiny appearance. It is made by adding lead oxide to the batch before melting. The designs or decorations are put on by means of grinding wheels after the glass has cooled.

Glass made with a high percentage of silica, some boric oxide, very little soda, and no lime is more resistant to heating and cooling and to chemical reagents. Such borosilicate glass is used under various trade names for kitchenware and for laboratory equipment. Articles made of it do not expand on heating nor contract on cooling nearly so much as if made of ordinary glass. For this reason, great strains are not set up in them as they are heated and cooled, and they are much less likely to break.

EXPERIMENT 289—Why Glass Cracks When Suddenly Chilled.

Heat a small piece of glass tubing in your alcohol lamp flame until it softens, and then plunge it quickly into water. The tube will crack due to unequal contraction on sudden cooling. Of course, when you plunge the hot tube into cold water the surface of the glass cools instantly, while the inside of the glass is still hot, and the unequal contraction which takes place when the glass is cooling causes it to crack.

Glass for special purposes may contain in addition to the usual constituents a great variety of others, for example, potassium, zinc, barium, lead, boron, phosphorus, arsenic, aluminum and antimony. Colored glass is made by adding to the molten batch suitable metal oxides or salts. The same substance may produce different colors, depending upon the amounts used and the method of adding it.

EXPERIMENT 290—Coloring Glass Light Green.

Make a small batch of soft glass in your spoon by heating intensely a mixture of sodium silicate, sodium carbonate and calcium oxide, as you did in the first experiment of this section, but add in addition a tiny crystal of Ferric Ammonium Sulfate (No. 21) about the size of a pinhead. You will find that your glass turns out to be greenish yellow.

If you look edgewise through a broken piece of cheap bottle glass, you will see that it is greenish. This is because of the iron contained as an impurity in the raw materials from which the glass was made.

EXPERIMENT 291—Coloring Glass Amber.

Make a glass mixture as before, adding a tiny amount of powdered charcoal. Notice the beautiful amber color which results.

EXPERIMENT 292—Coloring Glass Dark Blue.

Repeat the same process of glass-making, but this time use 1 drop of Cobalt Chloride (No. 27). Upon fusing this in your glass mix, you will have a glass of a dark blue color.

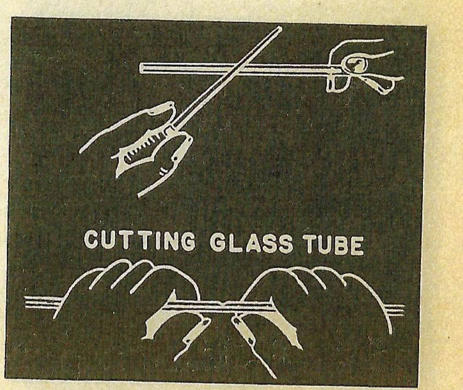
The value of glass as a building material is just beginning to be appreciated. Hollow glass blocks are used for the walls of windowless buildings, and beautiful decorative effects may be obtained, since the blocks can be made in any desired colors.

Fiber-glass is a fluffy, wool-like material which was developed for use in air filters for air conditioning installations. It is made by blowing jets of high-pressure steam against streams of melted glass. The individual glass fibers are only one-twentieth the diameter of a human hair, and a single pound of glass (the same amount used to make a beverage bottle) will make a single thread of more than 40,000 miles long. It is used in chemical plants for filtering corrosive liquids that would attack other filtering materials. Since it will not burn, it is valuable for insulating the walls of buildings against heat and cold. Its high strength and its resistance to fire suggest wide use in the manufacturing of rugs, curtains and electric wire insulation for reducing fire hazards in theatres, homes, factories, and ships at sea.

In working chemical experiments, it is often necessary to cut or bend glass tubing, and it is a great advantage to be able to do this well.

EXPERIMENT 293—How to Cut Glass Tubing.

Make a scratch with a three-cornered file in the place in which you wish to break the tube. Now hold the tube in your hands with your thumbs pressing upon the tube directly underneath the scratch, and then apply a pressure outward with your hands and break the tube just as you would break a stick. See the accompanying illustration. With a little practice, a clean and even fracture can be obtained.



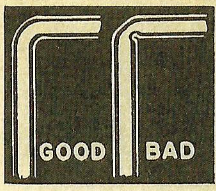
EXPERIMENT 294—Fire-Polishing Glass Tubing.

The cut ends of the piece of glass tubing have extremely sharp edges and are likely to cut stoppers and rubber tubes into which they are forced, to say nothing of your hands. It is good practice, therefore, to round off the sharp edges by fire-polishing. This is done by holding the end of the tube in the alcohol lamp flame until the glass softens and the edges become rounded. Observe that the hottest part of the flame is some distance above the wick. Be sure to rotate the tube slowly while you heat it so that it is heated evenly all around.

EXPERIMENT 295—How to Bend Glass Tubing.

Place the part of the glass tube where the curve is desired in the flame of your alcohol lamp. The whole surface of the tube must be heated, so it is necessary to turn it in the flame to heat all sides. It is also best to use as wide a flame as possible so that the tube will be heated for a considerable length. This makes it easier to obtain a gradual bend and one that will not break easily. When the tube becomes hot and soft, remove it from the flame and bend it to the desired angle by simply applying a little pressure to the ends and bending it in much the same manner as you would a piece of wire.

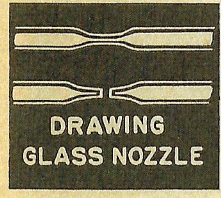
Unless you heat the tube for a considerable distance your bend will kink and may not allow the proper passage of gases or liquids through the bent tube. This type of bend will break if slightly twisted. The illustration shows the RIGHT and WRONG way to bend glass tubing.



EXPERIMENT 296—How to Make Glass Nozzles.

A glass nozzle or tip is very convenient for use in burning small quantities of gas, as the gas coming out of a very small opening will burn better than when coming out of a large opening.

To make a glass nozzle, take a piece of tubing, at least 3 or 4 inches long and heat one part of it, turning it in the flame so as to heat all sides equally. When the glass begins to soften remove it from the flame and pull the tubing out, that is, pull on each end of the tubing and you will find that the place where the glass is heated will stretch out, becoming smaller and smaller until finally it is almost the size of a thread. When the tubing becomes cool, break off at any desired point, depending on the size of the opening which you want in the end of the nozzle. When breaking the thread, it is best to use a file scratch in order to be sure of a good, clean break. Of course, when you make a nozzle in this way you really have two nozzles, but you are quite likely to spoil one before you are through.


DRAWING GLASS NOZZLE

EXPERIMENT 297—How to Close the End of a Glass Tube.

Hold the end of the piece of tubing which you wish to seal or close in the hottest part of the alcohol lamp flame. Turn it between the thumb and fingers and continue to heat and you will see the end gradually coming together, until finally, the tube is entirely closed. Another way of closing a piece of glass tubing which is very much better if the tube is of large diameter, is to heat the tube all around as though you were going to bend it and when the glass becomes soft, pull from both ends and draw the glass out into a fine thread. Now break the two pieces apart and hold the end in the flame just long enough to form a nice round bottom where the glass has been pulled out.

EXPERIMENT 298—Joining Glass Tubes.

Take two pieces of tubing and heat the ends of each in the alcohol lamp flame at the same time until the edges start to round and curve inward. Then quickly press the two sections together. To make a smooth joint, it is necessary to reheat the place where the tubes are joined and work the glass together. This can be done best by closing one end of the joined tubes and blowing in the other very gently while the joint is hot and soft. By repeating this process a number of times you will find that the joint can be worked down until it is smooth, even, and strong.

EXPERIMENT 299—Etching Glass.

Obtain some very strong liquid glue and using a brush, paint a thick coat over the outside and inside of a tumbler. Set the glass away for about a day until the glue becomes dry and then put the tumbler in a warm place over a heater or near a stove, and allow it to remain there for about a day longer.

Now remove the glass from the warm place and put it where it is cool. About this time the glue will begin to crack and in doing so, will break off small chips of glass from the surface of the tumbler. After the glass has remained in a cool place for a little while, wash it with warm water to remove all traces of glue, and you will find very peculiar designs etched on the surface. If this experiment does not work properly with one coating of glue, apply two or more coats before putting it in a warm place. A very strong glue must be used.

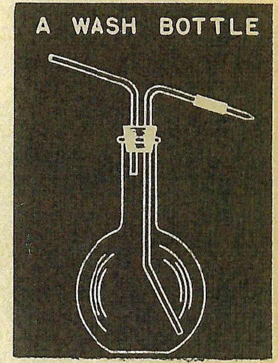
EXPERIMENT 300—Frosting Glass.

Obtain a small quantity of Epsom salts, which is the same as magnesium sulfate. Dissolve 3 teaspoonfuls of Epsom salts in a glass of water and add about 1 teaspoonful of liquid glue, stirring the mixture thoroughly. Now apply this solution with a brush to the glass which you wish frosted. When the solution dries on the glass it will produce a splendid frosted appearance. This solution can be used to frost electric light bulbs, windows and many other articles.

EXPERIMENT 301—Making a Wash Bottle.

If you do not have running water in your laboratory, here is a piece of equipment which is easily made and will prove indispensable. To make a wash bottle, bend two pieces of glass tubing as shown in the illustration. You will also need a two-holed stopper, drawn glass nozzle as made in one of the experiments above, a small piece of rubber tubing and a Florence flask or bottle. Insert the two bent tubes through the stopper, first being sure that the ends are properly fire-polished, and attach the nozzle to the longer tube by means of the short piece of rubber tubing. Fill the flask ¾ full of water and place the cork assembly in the flask as shown. By blowing in the end of the shorter glass tube, you can direct a steady stream of water from the nozzle into test tubes, beakers or whatever you desire.

A WASH BOTTLE



PAINTS, PIGMENTS AND COLORS

Since the dawn of history, the human race has been attracted by bright colors. Various kinds of paints were used thousands of years ago. Archeologists have uncovered Egyptian tombs in which were found beautiful colored paintings that have withstood the ravages of time. The American Indians used natural, mineral and vegetable colors to adorn themselves and their belongings. Today we are surrounded by colors on all sides and the use of color in our homes, factories and daily surroundings has almost become a science.

Coloring material used for protective and decorative purposes can be separated into two general classifications—paints, which consist of insoluble pigments mixed with a suitable vehicle or liquid—and stains, or dyes, which consist of a soluble color dissolved in a suitable liquid medium.

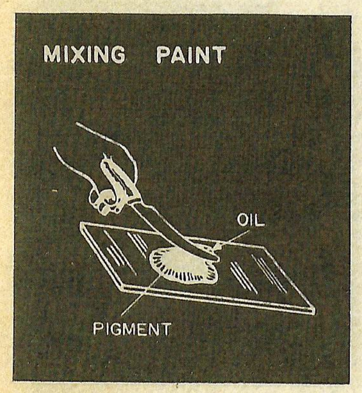
Paints can be further divided according to the vehicle, or liquid, used to carry the pigment. In general, these may be natural oils, such as linseed oil; synthetics, such as lacquers or liquid plastics; and finally water. Stains or dyes, may consist of organic solvents, such as turpentine, naphtha or alcohol, or the solvent may be water, depending upon the use of the stain.

Oil paints are by far the most important of our coloring materials. These paints are used for outside work where protection of the surface is of importance, as well as the beauty of the color. Oil paints usually consist of linseed oil into which the pigment or color is very finely ground. Turpentine, or other thinner, is added to bring the paint to the proper consistency for brushing or spraying, and in some cases, dryers are added to hasten the drying process. Linseed oil, when exposed to the air, is oxidized, thus forming a hard, flexible, lasting film on almost any rigid surface, such as wood, metal or brick. In the following experiments you will prepare some different colored pigments and grind them with linseed oil and turpentine to demonstrate the preparation of this class of paints.

EXPERIMENT 302—Making a Brown Pigment.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube ⅓ full of water. In another test tube ⅓ full of water, dissolve 4 measures of Sodium Carbonate (No. 4). Pour the first solution into a second, shake the tube, and you will obtain a thick brown precipitate of ferric hydroxide. Let this precipitate settle, pour off the clear liquid and then scrape the precipitate into your spoon. Warm this precipitate in your spoon over a flame until it is thoroughly dried and no more moisture is given off. You now have brown oxide of iron, which is used as a pigment in brown paint.

EXPERIMENT 303—Brown Oil Paint.



Place the pigment which you prepared in the preceding experiment on a piece of clean glass, add a few drops of linseed oil and thoroughly mix the oil and the pigment with a flexible knife blade or with the back of a teaspoon. It will require quite a little mixing and rubbing to make a smooth oil paint. If the paint seems a little thick when you get it thoroughly mixed, add a few drops of turpentine, which will thin it so that it will brush out more smoothly.

EXPERIMENT 304—Making a Light Blue Oil Paint.

Place about 6 measures of Azurite (No. 39) on a glass plate, add a few drops of linseed oil and mix and grind this as before, adding a little turpentine, if necessary. This will make a very pretty light blue oil paint, sometimes called "mountain blue."

EXPERIMENT 305—Making Black Oil Paint.

See if you can obtain some lampblack or soot. If you or one of your neighbors uses fuel oil for heat, it should be easy to obtain a little soot from the inside of the furnace.

Place a small pile of soot on a glass slab, add a little linseed oil and again grind the mixture to the proper consistency. Lampblack is often used as a pigment in black oil paint.

Lacquers and synthetic enamels are often made from cellulose acetate or other synthetic plastic material dissolved in a suitable organic solvent or mixture of solvents. If no pigment is added, we have what is known as a clear lacquer. If a pigment is ground into the lacquer, we have the various colored lacquers with which we are familiar.

Synthetic plastic enamels are often made which require the application of heat to set them. This is not a simple drying, but brings about a change in the molecular structure of the plastic material. Automobiles are often finished with these materials.

EXPERIMENT 306—Making a Brown Lacquer.

Prepare some brown oxide pigment as in a previous experiment and when it is thoroughly dried, mix it on a glass slab with a few drops of clear liquid nail polish. If your lacquer is too thick when mixed, add a few drops of nail polish remover. Note that your lacquer, when applied to a surface, dries almost immediately. This is one of the big advantages of lacquer, as oil paints dry quite slowly due to the time required for the linseed oil to oxidize.

EXPERIMENT 307—Making Black Lacquer.

Proceed as in the above experiment, but use some lampblack instead of the brown iron oxide. This time you will obtain black lacquer, which should cover quite well and be very durable.

Water paints have a wide field of usefulness, principally of course, for indoor work where they will not be subject to the action of the weather. An exception is whitewash which, even though mixed with water, will stand up outdoors for quite a long time.

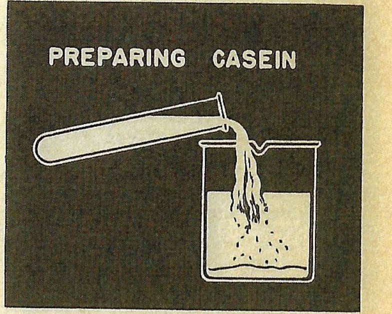
EXPERIMENT 308—Making Whitewash.

Put 5 measures of Calcium Oxide (No. 20) in a cup and add 5 or 6 drops of water. Stir this mixture until it forms a smooth thin paste. This is whitewash. Whitewash is used extensively for painting fences and rough wooden buildings. The particles of lime fill the pores of the wood and since the lime is not soluble in water, the coating will last for quite a long period of time. If desired, a little glue can be added to the whitewash to make it stick better when applied to a surface.

Paints which will stand up under fairly severe conditions can be made, using a solution of casein in water together with any desired pigment to add color. Freshly prepared casein is readily taken up by water forming a colloidal suspension or solution, especially if a little lime or other alkaline substance is present. After drying, however, the casein becomes insoluble, so that a wall painted with casein paint can be washed after the paint has thoroughly dried.

EXPERIMENT 309—Preparing Casein.

Fill a tumbler about ¼ full of skim milk. In a test tube half full of water dissolve 10 measures of Sodium Bisulfate (No. 7). Add this solution, a few drops at a time, to the skim milk, stirring the milk continually with a spoon. After the sodium bisulfate has all been added to the milk, you will notice a gelatin-like precipitate being formed in the milk and upon allowing it to stand for a while, the precipitate becomes heavy. This precipitate consists principally of casein. Carefully pour off the liquid, saving the casein in your glass for a future experiment.



EXPERIMENT 310—Making Brown Casein Paint.

Prepare some brown iron oxide pigment as in a previous experiment and carefully mix this pigment together with ½ measure of Calcium Oxide (No. 20) and some of the casein which you have just prepared. You will find that the casein is very soft and that the brown pigment can be ground into it, making a smooth paint. This paint can be thinned with water, if desired. Once your casein has dried, it cannot be again mixed with water as it has become insoluble.

EXPERIMENT 311—Making Black Casein Paint.

Proceed as in the above experiment, but use lampblack as a pigment instead of brown ferric oxide.

EXPERIMENT 312—Making Light Blue Casein Paint.

Mix 5 or 6 measures of Azurite (No. 39) and ½ measure of Calcium Oxide (No. 20) with a small amount of your soft casein. This will make a very pretty blue paint, which can be thinned to the desired consistency with water.

Poster colors consist of pigments ground with a clear mucilage or glue. These colors are quite brilliant and in practice, they are usually applied to white or colored cardboard. They are sometimes called show-card colors.

EXPERIMENT 313—Purple Lake Poster Color.

In this and the following experiments you will need a small bottle of clear mucilage with which to mix your colors. Clear mucilage can be obtained in almost any five and ten cent store or drugstore.

Put 2 measures of Logwood (No. 29) in a test tube half full of water and heat the tube for 3 or 4 minutes to extract the color. When the water in the tube has become a deep brown color, add 1 measure of Aluminum Sulfate (No. 12). Shake the tube thoroughly and carefully pour the clear red solution into another test tube, leaving behind the chips of logwood and any other sediment. Add to this clear red solution 1 measure of Sodium Carbonate (No. 4) and again shake the tube. The result will be a deep purple precipitate. Let this tube stand overnight until the precipitate settles, then pour off the clear liquid and scrape the purple precipitate into a small cup or onto a flat piece of glass. This purple pigment, when mixed with clear mucilage, will make a very pretty purple poster paint.

In this and several following experiments much time can be saved by filtering the liquid to separate the precipitate. See the instructions for filtering under "The Use of Chemical Apparatus."

EXPERIMENT 314—Making Black Lake Poster Color.

A black pigment, very similar to the purple pigment which you just made, can also be prepared from logwood. Heat 2 measures of logwood in a test tube half full of water for several minutes until you extract the color from the logwood. Add 1 measure of Ferric Ammonium Sulfate (No. 21) and shake the tube thoroughly. Pour off the black solution which is formed, being careful to leave behind the chips of logwood and any other sediment. Now add to this black solution, 1 measure of Sodium Carbonate (No. 4). Shake thoroughly and then let the black precipitate settle, so that you can pour off the clear liquid on top. This black pigment, when mixed with clear mucilage, makes an excellent black poster color.

EXPERIMENT 315—Light Blue Poster Color.

Mix 3 or 4 measures of Azurite (No. 39) with a few drops of mucilage. The result will be a pretty light blue poster color.

EXPERIMENT 316—Brown Poster Color.

Dissolve 2 measures of Ferric Ammonium Sulfate (No. 21) in a test tube ⅓ full of water. Dissolve 4 measures of Sodium Carbonate (No. 4) in another test tube ⅓ full of water. Pour the first solution into the second and you will obtain a thick brown precipitate which consists of ferric hydroxide. Let this precipitate settle, pour off the clear liquid from the top, and then scrape the precipitate from the test tube into a cup or onto a glass plate. This brown pigment, when mixed with clear mucilage makes a rich brown poster color.

EXPERIMENT 317—Green Poster Color.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ⅓ full of water. Now add 2 or 3 drops of Cobalt Chloride Solution (No. 27). Shake the tube and you will obtain a thick green precipitate of cobalt ferrocyanide. Allow this precipitate to settle, pour off the clear liquid on the top and scrape the green pigment onto a glass plate and mix with clear mucilage to form green poster color.

EXPERIMENT 318—Red-Brown Poster Color.

Put 2 measures of Azurite (No. 39) and 4 measures of Sodium Bisulfate (No. 7) in a test tube and fill it half full of water. Warm the tube for a few moments until the solution becomes a clear blue color and all the solids dissolve. In another test tube, dissolve 2 measures of Sodium Ferrocyanide (No. 6), shaking to dissolve all of the solid. Now pour the contents of the first tube into the second and you will obtain a red-brown precipitate of copper ferrocyanide. By allowing this precipitate to settle, pouring off the clear liquid and then scraping the precipitate onto a glass plate and mixing it with clear mucilage, you will obtain a reddish-brown poster color.

EXPERIMENT 319—Lampblack Poster Paint.

Proceed as in the previous experiments but use lampblack as a pigment and mix thoroughly with the clear mucilage. This will give a dead black poster paint.

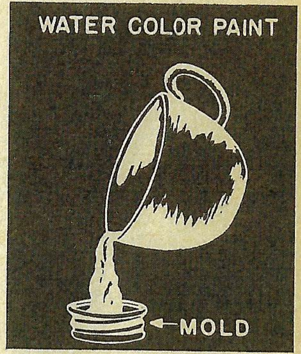
EXPERIMENT 320—White Poster Paint.

Zinc oxide makes one of the best white pigments for poster paint. Precipitated chalk or chalk dust can also be used as a pigment for white poster paint.

Artists' water color paints usually consist of a pigment or dye mixed with glucose, or a similar water soluble base. The glucose, or other base, is often heated while the color is added and upon cooling, the base then solidifies into a solid cake. Ordinary corn syrup is glucose and if you have some of this available, you can easily make a few water color paints.

EXPERIMENT 321—Blue Water Color Paint.

Dissolve 2 measures of Sodium Ferrocyanide (No. 6) in a test tube ¼ full of water. In another test tube ¼ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). Pour the second solution into the first and a deep blue color will result. Put a tablespoon of corn syrup into an old china cup and float the cup on a pan of hot water until the corn syrup becomes warm and fairly fluid. Now add a little of the blue color which you have prepared and mix it thoroughly with the warm corn syrup. This mixture can be poured into small molds and after it stands for a few hours, the cakes of paint will be hard just like regular water colors.



EXPERIMENT 322—Another Blue Water Color Paint.

Cut a small piece of Dye Wafer (No. 8) about ¼ inch wide by ½ inch long. Put this in a test tube ¼ full of water and shake the tube for a few minutes to extract the dye. Now warm some corn syrup as in the preceding experiment and mix the blue liquid from the dye wafer with the glucose. Pour it into molds and allow it to solidify as before.

EXPERIMENT 323—Black Water Color Paint.

Dissolve 1 measure of Tannic Acid (No. 15) in a test tube ¼ full of water. In another test tube ¼ full of water dissolve 1 measure of Ferric Ammonium Sulfate (No. 21). When the contents of both test tubes are dissolved, pour the tannic acid solution into the ferric ammonium sulfate solution and you will obtain an intense black color. Mix some of this black color with warm corn syrup and pour this mixture into molds to solidify. You will have cakes of black water color paint.

EXPERIMENT 324—Yellow Water Color Paint.

Put 2 measures of Logwood (No. 29) in a test tube ⅓ full of water and heat the tube for 3 or 4 minutes until the water has become colored by the logwood. Add ½ measure of Sodium Bisulfate (No. 7). Now add some of this yellow dye to warm corn syrup as before and the result will be a yellow water color paint.

EXPERIMENT 325—Other Colors of Water Color Paint.

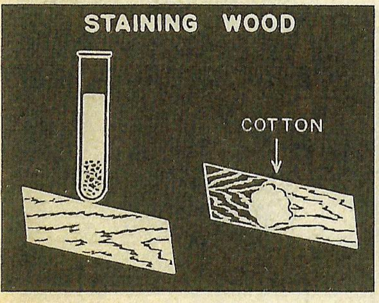
There is no limit to the colors that you can make up in these paints if you have water soluble dyes of various colors and shades. Perhaps you may have some packages of Easter egg dyes or dyes used in coloring fabrics; even colored inks can be used. Follow the directions in the preceding experiments in making up various other colors of water color paints.

Stains consist of soluble dyes or coloring material dissolved in a liquid. Stains for use out of doors where they are exposed to the weather are usually made with naphtha, turpentine, lacquer thinner or other organic solvent. Stains for use indoors can be made with water, although for woodwork, floors and other places where washing is necessary, the organic solvents are usually best.

Stains cannot be used on surfaces into which the liquid cannot penetrate. Thus, the principal use of stains is on wood. Woodwork is often stained to give the desired color or tone and then covered with a clear varnish or lacquer, which gives a glossy finish.

EXPERIMENT 326—Brown Lacquer Thinner Stain.

If you can obtain some nail polish remover, fill a test tube about ¼ full and add 2 measures of Logwood (No. 29). Use a cork, or crumple up a small piece of paper to stopper the mouth of the test tube and set it aside for several hours. The lacquer thinner will gradually take up color from the logwood forming, after some time, a very pretty yellowish brown stain that can be applied to any white wood with good effect.



EXPERIMENT 327—Brown Alcohol Stain.

Fill a test tube ¼ full of alcohol (rubbing alcohol will be satisfactory). Add 2 measures of Logwood (No. 29) and let this soak in the alcohol for an hour or so. This will form a brown stain which penetrates evenly into the pores of wood and makes a very attractive finish.

EXPERIMENT 328—Blue Lacquer Thinner Stain.

Cut a small piece of Dye Wafer (No. 8) about ¼″ wide by ½″ long and place this in a test tube ¼ full of nail polish remover. The lacquer thinner will gradually dissolve the blue dye and if left standing for some time, will produce a blue stain.

EXPERIMENT 329—Blue Alcohol Stain.

A small piece of Dye Wafer (No. 8) can be used with alcohol to produce a blue stain. The alcohol will dissolve the color better than the lacquer thinner and will make a more intense stain.

EXPERIMENT 330—Blue Water Stain.

Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ⅓ full of water. In another test tube ⅓ full of water, dissolve 1 measure of Sodium Ferrocyanide (No. 6). Pour the second solution into the first and you will obtain a blue color which can be used for staining wood or other absorbent material.

EXPERIMENT 331—Black Water Stain.

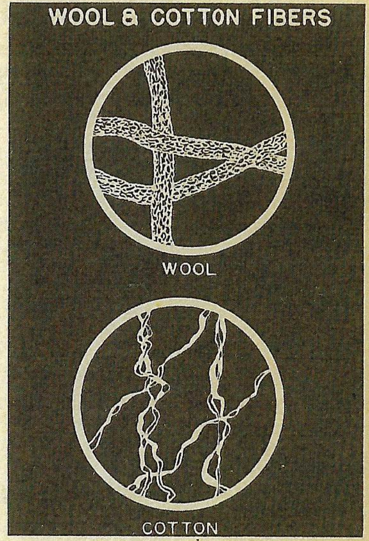
Dissolve 1 measure of Ferric Ammonium Sulfate (No. 21) in a test tube ¼ full of water. Dissolve 1 measure of Tannic Acid (No. 15) in another test tube ¼ full of water. When these substances are all dissolved, pour the tannic acid solution into the ferric ammonium sulfate solution and you will obtain a black color which can be used as a water stain.

EXPERIMENT 332—Yellow Logwood Stain.

Put 2 measures of Logwood (No. 29) in a test tube ¼ full of water. Warm the liquid for 3 or 4 minutes and add ½ measure of Sodium Bisulfate (No. 7). This makes an excellent stain for use on a light colored wood.

TEXTILE FIBERS

The textile industry is of great importance in our modern world. The business of growing or making the fibers, spinning them into thread, weaving the thread into cloth and making the cloth into useful articles employs a great many people. The output of the textile industry touches our daily life in many ways. Our clothing, rugs, bedding, curtains, the cords in automobile tires, the insulation of electric wires, cords and ropes of all kinds from cotton string to the great hawsers of ships, the binding of books, and dozens of other indispensable articles are products of the textile industry.



There are three general classes of textile fibers:

1—Fibers obtained from animals, such as wool, silk and fur.

2—Fibers obtained from plants, such as cotton, flax, jute, hemp and manila.

3—Fibers made from synthetic products, such as rayon, estron, nylon, casein and glass.

NATURAL FIBERS

Vegetable fibers, such as cotton and linen, consist chiefly of cellulose and are decomposed by strong acids. Animal fibers, such as silk and wool, consist largely of compounds containing nitrogen, carbon, hydrogen and sometimes sulfur. These withstand the action of acids but are readily destroyed by alkalies.

Many people can tell whether a fiber is silk, wool or cotton by feeling it and examining it closely. It is impossible, however, to detect certain mixtures of fibers in cloth by these methods, and so standard chemical tests have been devised by means of which the exact composition of cloth can be determined. Examination under the microscope is one of the best ways to distinguish these fibers.

EXPERIMENT 333—Testing Fabrics for Animal Fibers.

Animal fibers, such as silk and wool, burn slowly, giving an odor similar to burnt hair or horn, and form round granules of carbon at the point of the thread where combustion ceases.

EXPERIMENT 334—Testing Fabrics for Vegetable Fibers.

Vegetable fibers, such as cotton, artificial silk, and linen, burn rapidly, form a thin, white ash and give off only a faint smell, similar to burnt paper.

Take threads from various kinds of fabrics and burn them, watching the results closely. This is the simplest test for determining various fabrics.

EXPERIMENT 335—Testing Fabrics for Pure Linen Fibers.

Dissolve ½ measure of Sodium Bisulfate (No. 7) in a test tube half full of water. Put a small scrap of linen in this solution and boil it for a few minutes. Remove the linen, wash it carefully in water and then dry it.

Now fringe out the edges of the piece of cloth so as to expose the fibers and moisten one corner with a drop of glycerine. Press the cloth between blotters and then lay it against a dark background. The linen fiber will be translucent and if any other kind of fibers are mixed with the linen, they will appear opaque.

EXPERIMENT 336—The Alkali Test for Animal Fibers.

Wool and natural silk are dissolved by boiling them in sodium hydroxide solution while cotton, linen and artificial silk are not affected.

Prepare a solution of sodium hydroxide by dissolving 2 measures of Calcium Oxide (No. 20) and 2 measures of Sodium Carbonate (No. 4) in a test tube half full of water. Boil this solution for a few minutes, and then put in a small scrap of wool or silk and boil for a few minutes more. Be careful that the liquid does not spurt out of the tube into your face. Notice the effect on the fibers. In case a fabric is a mixture of wool and cotton, the wool fibers will be dissolved and the cotton fibers will remain.

SYNTHETIC FIBERS

At present the most important synthetic textile fibers are rayon, estron and nylon. These fibers have almost replaced natural silk and they are undoubtedly making inroads on the other natural fibers.

The basic material for both rayon and estron is cellulose. The cellulose may be in the form of specially purified wood pulp, or of cotton linters, that is, the short fibers removed from the cotton seeds by a second ginning.

Most of our rayon is made by the so-called viscose method, in which the purified cellulose is treated with sodium hydroxide solution and with carbon disulfide. The reaction which occurs forms a yellow spongy mass of cellulose xanthate. When more sodium hydroxide is added, this becomes a syrupy liquid, which is forced through tiny holes in a metal plate into a bath of sulfuric acid. When the filament of liquid comes in contact with the sulfuric acid, the sodium hydroxide in it is neutralized and the cellulose xanthate is decomposed, giving back the original cellulose, but in the form of fine, uniform fibers of unlimited length. These may be twisted into thread.

Two less-used processes, the cellulose nitrate and the cuprammonium, differ only in the means used to get the cellulose into liquid form. In both of them the liquid is converted into cellulose at the end of the process.

Cellulose acetate, formerly called acetate rayon, now known as estron, differs from that made by the viscose process in that it is not cellulose. The starting material is cellulose, just as in the other process. Treatment with strong acetic acid converts the cellulose into cellulose acetate. This is dissolved in acetone, squirted through fine nozzles to form it into fibers and the acetone removed. It remains, however, in the form of cellulose acetate instead of being converted back into cellulose.

Nylon was first produced commercially in the late 1930's. In chemical composition, it differs greatly from rayon and estron since it contains nitrogen in addition to carbon, hydrogen and oxygen. Chemically, it is more like natural silk. By a complicated series of reactions, suitable compounds are obtained from basic raw materials — coal, air and water. These compounds are melted together and slowly react with one another, forming long chains of atoms. The melted material is forced through tiny holes just as in making rayon and it then hardens in the form of filaments which, after a stretching process, are spun into thread.

Other synthetic fibers include polyvinylidene chloride under the trade name of saran, casein fibers made by extruding casein into a solution of formaldehyde, and polyacrylonitrile fiber, which is being produced under the trade name of orlon. In the section on glass, we mentioned the very fine glass fibers which it is now possible to make. They are finding increasing application in the textile field.

The various synthetic fibers have quite different physical characteristics, which makes them adaptable for different purposes.

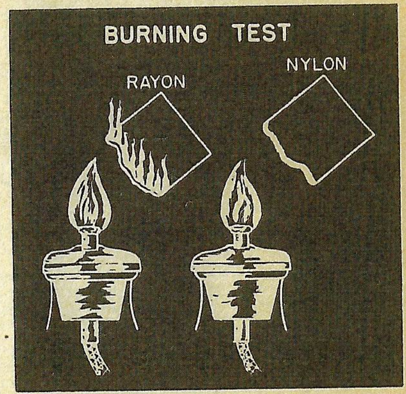
Viscose rayon is very susceptible to moisture changes and material made from these fibers will swell and absorb up to 75% of its weight of water. On the other hand, viscose rayon has high tensile strength and is quite resistant to heat, which makes it ideal for use in automobile tire cord and for many other purposes.

Cellulose acetate, estron, has a considerably lower moisture absorption and is faster drying. It is, however, much more sensitive to heat, melting at a considerably lower temperature.

Nylon combines high tensile strength, low moisture absorption and a relatively high melting point, which makes it suitable for purposes where great strength and resistance are required.

EXPERIMENT 337—The Burning Test for Synthetic Fibers.

Cut a small square of rayon or estron and of nylon from worn-out clothing or rags. Hold the edge of each square in a flame for a moment. If your sample is nylon, it will cease to burn as soon as removed from the flame. If it is either rayon or estron, it will continue to burn slowly when removed from the flame. If it is estron (cellulose acetate), the smoke which comes from the flame will have a characteristic sharp odor. If it is rayon (viscose), it will continue to burn slowly, the smoke smelling somewhat like burning paper.



EXPERIMENT 338—Melting Point of Synthetic Fibers.

Test small samples or rayon, estron and nylon under a hot iron. The cellulose acetate, or estron, will melt first and the viscose and nylon will stand a considerably higher temperature.

EXPERIMENT 339—Solubility of Synthetic Fibers.

Rayon, estron and nylon are not soluble in ordinary petroleum derivatives, such as naphtha or gasoline. They are all soluble, however, in other organic solvents such as lacquer thinner. This may be demonstrated by placing small squares of cloth made from different synthetic fibers in a test tube and adding a small amount of nail polish remover. Thus, care must be taken not to moisten cloth made of these fibers with liquids such as lacquer thinners.

In many ways the synthetic fibers are superior to our natural fibers, but on the other hand, natural fibers have certain important physical characteristics which so far we have been unable to duplicate in the synthetic fibers.

DYES

In the process of dyeing, the coloring matter is dissolved in water or other solvents and the material to be dyed passed through the solution.

Many dyes unite directly with the fiber of the cloth, forming, in some cases, colors which are insoluble and will withstand washing; in other cases, the colors formed in this manner are not entirely "fast" and will "run" to some extent when the cloth is washed.

Since there are many different kinds of fibers used in making textiles, we are not surprised that the same dye gives different shades and intensity